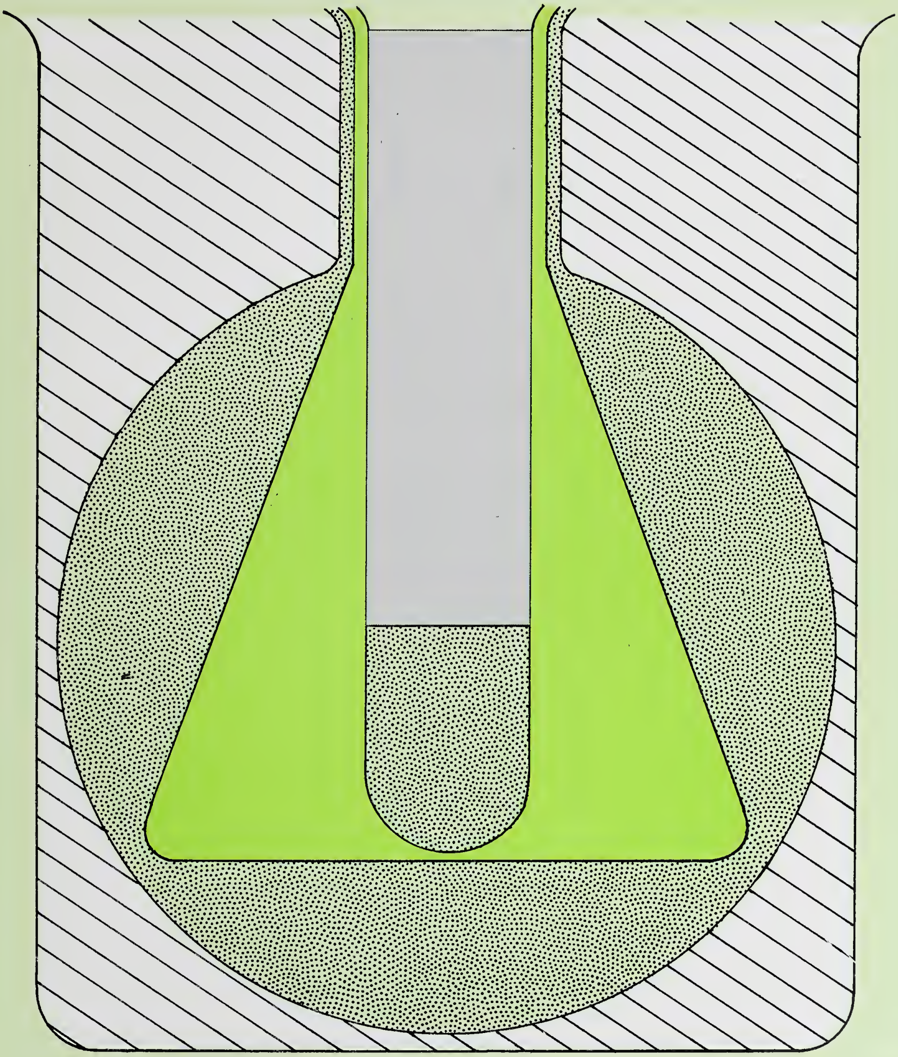
CHEMISTRY



10

ALBERTA CORRESPONDENCE SCHOOL
ALBERTA EDUCATION

Chemistry 10

Lessons 1-12

Distance Learning

Alberta
EDUCATION

ISBN No. 0-7741-0015-X

Chemistry 10 Students

You will need this material for experiments in Chemistry 10. Until the material arrives, you may want to work on the theory part of Lessons One and Two.

The materials are only loaned to students living in Alberta who have purchased the Chemistry 10 laboratory kit from the Alberta Correspondence School. To obtain this equipment, complete the application form on the reverse of this page and send it to:

Alberta Correspondence School

Box 4000

Barrhead, Alberta

T0G 2P0

After the experiments are done, immediately return the materials to the Alberta Correspondence School in the condition you received them.

Students living outside Alberta are responsible for obtaining these supplies from local schools or industries. As well the equipment can be purchased from scientific supply companies such as Boreal Laboratories Ltd., 1820 Mattana Avenue, Mississauga, Ontario, L4X 1K6, or, Central Scientific Company, 2200 S. Sheridan Way, Mississauga, Ontario.

Check local outlets before you purchase any of these items.

APPLICATION FOR LOAN OF A BEAM BALANCE, A SET OF WEIGHTS, A THERMOMETER, AND A GRADUATED CYLINDER

Name: ______________________________ File Number: __________

Address: ______________________________

____

____

I have purchased a laboratory kit for Chemistry 10 and reside in the Province of Alberta. I am ready to use the beam balance and set of weights and a thermometer and graduated cylinder to perform the required laboratory experiments.

Please send me **on loan** the above items.

I agree to use the material carefully and to return it in good condition after the experiments are completed.

____
(Signature)

Note: The Alberta Correspondence School will **not lend** the items to a student with an address outside the Province of Alberta or to a student who has not purchased a Chemistry 10 laboratory kit.

N.B. Work on the theory part of the first few lessons until this material arrives.

Contents of Course

Lesson 4: Classification of Matter (continued)

- E. Mixtures, Pure Substances, Elements and Compounds
- F. Experiment 6 - Forces Between Magnetic Poles
- G. Energy and Chemical Changes
- H. Symbols
- I. Experiment 7 - Production of FeS
- J. Ways of Separating Matter
- K. Experiment 8 - Separation of Sand and Salt

Lesson 5: Atomic Theory

- A. Experiment 9 - Black Box
- B. Dalton's Model
- C. Gay Lussac and Avogadro
- D. Rutherford's Model
- E. Periodic Table

Lesson 6: Chemical Equations

- A. IUPAC Rules for Writing and Naming of Compounds
- B. Mole Concept

Lesson 7: Chemical Equations (continued)

- C. Writing Chemical Equations
  - (1) Qualitative
  - (2) Quantitative
- D. Balancing of Chemical Equations
  - (1) Using Moles of Atoms
  - (2) Using Atoms
  - (3) Conservation of Mass

Lesson 8: Chemical Equations (continued)

- E. Using Chemical Equations in Problem Solving
- F. Types of Chemical Reactions
  - (1) Synthesis Reaction
  - (2) Decomposition Reaction
  - (3) Rearrangement or Replacement Reaction
  - (4) Metal-Ion Reaction
  - (5) Ion-Ion Reaction
- G. Experiment 10 - A Quantitative Study of a Chemical Reaction.

Lessons 9 to 11: Elective: Metals

Lesson 12: Review

Materials Required for Each Experiment (Materials found in the kit are marked with a *.)

Experiment 1

- 10 cm^3 graduated cylinder (lent)
- balance (lent to students)
- iron nail
- 2 brass weights from balance
- * aluminum rod

Experiment 2

- * lead nitrate ($Pb(NO_3)_2$
- water
- * sodium iodide (NaI)
- 1 - 10 cm^3 graduated cylinder
- * 1 - 13 × 100 mm test tube
- * 2 pieces of filter paper
- * 1 beaker
- balance

Experiment 3

- 4 sugar cubes
- 4 glasses or paper cups
- hot and cold water
- watch or clock

Experiment 4

- * paradichlorobenzene
- * test tube
- candle
- * clamp
- thermometer (lent to student)
- water
- * 100 mL beaker

Experiment 5

- * 2 - 13 × 100 mm test tubes
- * small piece of Mg ribbon ($\approx$4 cm)
- quarter of a test tube of vinegar
- * wooden splint
- match
- * 100 mL beaker

Experiment 6

- * 2 ceramic magnets
- ruler
- balance

Experiment 7

* 100 mL beaker
* test tube clamp
balance
* 5 - 13 × 100 mm test tubes
* 0.1 M HCl (small bottle)
candle
matches
* sulfur (5 g)
* iron filings (5 g)
hot water
cold water
old rag
hammer

Experiment 8

* 100 mL beaker
another equally large container
* mixture of sand and salt (100 g)
spoon
balance
thermometer
oven or other source of heat

Experiment 9

* black box (see note in kit)

Experiment 10

* sodium iodide (NaI)
* lead nitrate ($Pb(NO_3)_2$
* 5 test tubes (13 × 100 mm)
tap water
10 cm^3 graduated cylinder
ruler (marked in cm)
* glass stirring rod
balance
* 100 mL beaker
labeling tape
test tube stand, (or a glass to hold 5 test tubes)
* eye dropper

Textbooks:

Keys to Chemistry by Ledbetter and Young

Laboratory Keys to Chemistry by Ledbetter and Young

General Information

At the present time, two courses have been approved for Alberta students in Chemistry 10 and Chemistry 20. They are the Alchem course, which has been written by a number of Edmonton Teachers, and the course based on the text Keys to Chemistry, which you are taking. Of course, there will be minor differences in what is taught in these two courses but the major concepts should be the same. A student should be able to go from one course to the next without too much difficulty. In covering certain concepts we have had to jump around a bit in the text and you may take Chapter 8 without having taken Chapter 7. However, we have tried to make it clear what you are expected to know and what you can leave for a future chemistry course.

The first eight lessons of this course are the core materials which every Alberta student must know. There is more leeway in what is taught in the last portion of the Chemistry 10 courses in general, therefore your final test will mainly stress the first eight lessons.

In this course, you will often be asked to do Practice Exercises and Self-Tests from the textbook. When you complete these exercises, check your own answers with those provided in the textbook. If you can't understand how a given answer was arrived at, feel free to ask about it, but you are not required to send in your answers to these self-checking exercises. All other lesson pages which have questions on them must be sent in for correction. Often, the questions in the lessons are very similar to the self-checking exercises.

Final Mark

Your final mark will be determined as follows:

Final Test	- 70% (emphasis on first 8 lessons)
*Lessons	- 30%

*The mark for the lessons will only be used when the student has scored a minimum of 40% on the final test. Students scoring less than 40% on the final test will receive that mark as their final grade.

In general, the assessment should not hurt your mark if you had been doing your own work. Often we get students who can do the moxt complex problems on the lessons and get straight A's, but can't do the simplest problems on the test. To get help is acceptable but you must understand how a given problem is solved so that you can do a similar problem on the test. It is to your advantage to perform all exercises by yourself to get the practice in doing chemistry.

IMPORTANT NOTE REGARDING FINAL EXAMS

Final exams in Chemistry 10 will be CLOSED BOOK. Students writing final exams will be allowed to bring to the exam only writing utensils, calculator and/or slide rule. A periodic table of elements and ions will be supplied with the exam. For regular school students, principals are in charge of the final exam and all questions pertaining to the exam arrangements must be directed to him. Non-classroom students must make their arrangements as suggested in the Information Bulletin.

Laboratory Work

The laboratory work is extremely important and the experiments furnish a basis for understanding the concepts of chemistry. When doing laboratory work, prepare yourself thoroughly beforehand so that you know exactly what you are trying to do and how you plan to do it.

Since you are taking this course by correspondence, you will probably be using the lab kit (unless you can use a school lab) which will enable you to do all the required experiments reasonably well. Even with a well equipped school lab, we will just expect you to do what a student can do who has our lab kit. If an exception should arise, as in Experiments 6 and 7, further instructions will be given.

Since it is impossible to send out all materials which a good school lab has, our procedures are somewhat different from those given in the lab manual. You are advised to read through each experiment in the lab manual since it gives helpful information. This will enable you to understand each lab better. Due to limited lab supplies, do what is indicated in the lesson notes rather than the lab manual and answer questions which are asked for in the lessons.

CAUTION: Do not taste any of the chemicals. Keep them out of the reach of children and away from pets. If a chemical spills on your work surface, your skin, or your clothing, wipe it immediately and rinse with plenty of cool water. Dispose of all chemicals in a safe manner after each experiment is finished.

Policy on Handling Chemicals

The following information is available to all students enrolled in courses offered by the Alberta Correspondence School which make use of chemicals:

1. A list of chemicals accompanies each laboratory kit that students purchase.

2. A warning notice is provided with each laboratory kit containing chemicals. This notice states that "all chemicals must be regarded as poisonous." On this notice students are informed to:

   a. store all chemicals out of reach of children.

   b. use the chemicals as instructed in the lesson notes.

   c. dispose of these chemicals promptly after use.

3. The lesson material that deals with handling chemicals contains information as to what students need to do if any of these chemicals are spilled on the work surface or on the skin and clothing.

   In general, if chemicals are spilled on the work surface soak them up with a sponge. Wear safety gloves while doing this procedure. Then dilute with excess water and rinse the chemicals down the drain. Wash the sponge carefully making sure the chemical residues have been washed away. The chemicals contained in each laboratory kit are not likely to damage the work surface (i.e. counter, table); however, it is recommended that students use a thick layer of papertowel or a rubber mat to protect the work surface from any spills that may occur.

   If chemicals come in contact with clothing, remove the clothing immediately and rinse thoroughly with excess water. Wear protective plastic gloves while doing this procedure.

   If the students' hands come into contact with chemicals they are to be washed thoroughly with water and then soap and water.

4. Information regarding the method of disposing of chemicals are included in the introduction of Chemistry 10 as well as within the lessons. In general, liquids and solids are to be diluted with water and flushed down the drain with large quantities of water. These solids and liquids can also be flushed down the toilet where they are automatically diluted. A few of the chemicals found in the lab kit need to be disposed in a special manner. In these particular cases extra chemicals will be provided in the laboratory kit to safely dispose of these chemicals.

5. It is recommended that students wear plastic gloves, some eye protection (goggles), and some protective clothing (i.e. lab coat or smock) when dealing with chemicals.

6. Washing procedure if acids and/or alkali (bases) get into the eyes:

   a. Acids: wash eyes with cold water for 20 min. and seek immediate medical attention.

   b. Bases (Alkali): wash eyes with cold water for 25 min. and seek medical attention immediately.

7. It is imperative that students do all their lab work near the sink or where water is available.

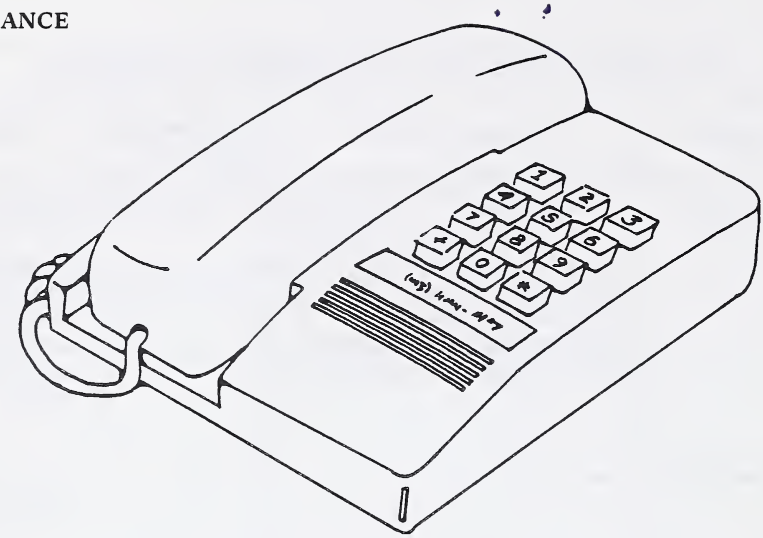
8. Students must remember to wash their hands thoroughly after each experiment.

Information Regarding the Disposal of Chemicals

Chemistry 10	Chemical	Description and Use	Is this chemical hazardous?	Method of Disposal
Experiment 5 Experiment 10	Lead Nitrate $Pb(NO_3)_2$	- White crystals - formed by reaction of lead oxide and nitric acid - decomposes on heating, this leaves lead oxide residue. - used to furnish a soluble lead salt. - used in photographic emulsions.	Yes	- make a solution of lead nitrate and mix it with a solution of sodium silicate (Na_2SiO_2). After adding the solutions together let them stand for awile. Take the entire solution and flush it down the toilet where it is automatically diluted.
Experiment 5	Sodium Iodide	- salt	No	- Throw in the waste paper basket or flush down the toilet.
Experiment 2	Paradichlorobenzene	- white solid - an effective multipurpose pesticide. This compound is particularly useful against the sugarbeet weevil.	No	- Throw in the waste paper basket or flush down the toilet
Experiment 3	Magnesium Ribbon	- metal strip - not harmful - gives off heat and light when burned	No	- Throw in waste paper basket making sure it is not lit.
Experiment 7	Hydrochloric Acid (HCl)	- 0.1 M very small concentration of HCl - this concentration is not harmful.	No	- Dilute with water and throw it down the sink or flush down the toilet.
Experiment 7	Sulfur	- only 2g is used. - present in amino acids and hormones. - present in meat, eggs, cheese. - sulfur is present in the human body in trace amounts.	No	- Flush down the toilet
Experiment 7	Iron Filings	- thin metal shavings	No	- throw in the waste paper basket.
Experiment 7	Iron Sulfide	- present in protein rich food	No	- flush down the toilet

SECTION C

SPECIAL ASSISTANCE



TOLL FREE

*** Method 1**

1. Look in your local telephone directory under the Government of Alberta for your local **RITE** number.

2. Dial the RITE number.

3. Ask for the Alberta Correspondence School in Barrhead.

NOTE: If there is no local RITE number use the following procedure:

*** Method 2**

1. Dial the Operator (0).

2. Ask for Zenith 22-333.

3. Ask for the Alberta Correspondence School in Barrhead.

* Can only be used while living in Alberta.

ADVANCE NOTICE CONCERNING TESTING AND COURSE EVALUATION

1. In order to be recommended for credits for Chemistry 10 you are required to write a supervised test set by the Alberta Correspondence School before registration expires. A portion of the final mark will be based on your course work, as evaluated by the teacher of the Alberta Correspondence School. If the final mark differs substantially from the year's work, the teacher will use discretion in balancing the composition of the marks in order to arrive at a fair assessment of achievement in the course. Appeal papers will be available to students who do not achieve a pass, and whose registration has not expired.

2. (a) **Classroom students**

Those who are in attendance in school in Alberta and who are supplementing their school program by taking one or more correspondence courses.

A student attending school does not submit an application for the final test. Test papers are sent automatically to the principal during the school year or at the end of August for writing during the first week of September for summer school students. **ELEVEN SATISFACTORY LESSONS** out of the twelve to be submitted, must be received by the Alberta Correspondence School before a test paper is mailed to the principal.

The principal is in charge of scheduling final tests and all questions about scheduling should be directed to the principal.

If a test is not written before expiry date of registration, the course is considered incomplete for the school year which the student registered.

(b) **Non-classroom students**

Those who are studying exclusively by correspondence, and are not registered in any subjects in the Alberta classroom.

To obtain course credits, non-classroom students must complete all required lessons and write the final test before expiry date of registation. Information about expiry dates is given in the Information Bulletin which a student receives before filing an application for a correspondence course.

The application for the final test is sent out when the corrected Lesson 6 is returned to the student and the student submits the application with Lesson 10. The test is sent out after **ELEVEN SATISFACTORY LESSONS**, out of the twelve to be submitted, have been recieved by the Alberta Correspondence School. However, all twelve lessons should be submitted before the test is written.

If a test is not written before registration expires, the course is considered incomplete for the school year during which the student registered.

NOTE: For the purpose of writing final tests, students who live outside Alberta come under the same regulations as those in category (b).

How Your Lessons Will Be Graded

Your standings on the lessons will be indicated by means of letter gradings on the following basis:

Letter Gradings	Range of Scale
A	80 - 100%
B	65 - 79%
C	50 - 64%
D	40 - 49%
F	0 - 39%

We wish you success and enjoyment in this course.

Periodic Table of Elements

TABLE OF COMPLEX IONS

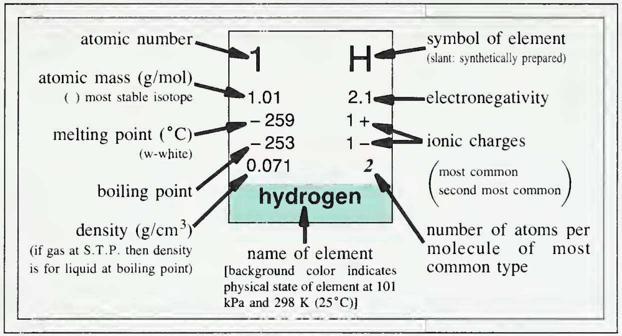
acetate	CH_3COO^-	glutamate	$C_5H_8NO_4^-$
ammonium	NH_4^+	hydrogen phosphate	HPO_4^{2-}
benzoate	$C_6H_5COO^-$	hydroxide	OH^-
bicarbonate (hydrogen carbonate)	HCO_3^-	hypochlorite	ClO^-
bisulfate (hydrogen sulfate)	HSO_4^-	iodate	IO_3^-
bisulfide (hydrogen sulfide)	HS^-	nitrate	NO_3^-
bisulfite (hydrogen sulfite)	HSO_3^-	nitrite	NO_2^-
borate	BO_3^{3-}	oxalate	$OOCCOO^{2-}$
bromate	BrO_3^-	perchlorate	ClO_4^-
carbonate	CO_3^{2-}	permanganate	MnO_4^-
chlorate	ClO_3^-	phosphate	PO_4^{3-}
chlorite	ClO_2^-	silicate	SiO_3^{2-}
chromate	CrO_4^{2-}	stearate	$C_{17}H_{35}COO^-$
cyanide	CN^-	sulfate	SO_4^{2-}
dichromate	$Cr_2O_7^{2-}$	sulfite	SO_3^{2-}
dihydrogen phosphate	$H_2PO_4^-$	tetraborate	$B_4O_7^{2-}$
ferricyanide	$Fe(CN)_6^{3-}$	thiocyanate	SCN^-
ferrocyanide	$Fe(CN)_6^{4-}$	thiosulfate	$S_2O_3^{2-}$
		tripolyphosphate	$P_3O_{10}^{5-}$

SOLUBILITY OF SOME IONIC COMPOUNDS IN WATER AT 298 K (25°C)

ION	Group IA		NH_4^+	NO_3^-	CH_3COO^-	Cl^- Br^- I^-	SO_4^{2-}	S^{2-}	OH^-	PO_4^{3-} SO_3^{2-} CO_3^{2-}
	alkali metal ions	H^+ (H_3O^+)								
SOLUBILITY >0.1 mol/L (VERY SOLUBLE)	All	All	All	All	Most	Most	Most	Group IA Group IIA NH_4^+	Group IA NH_4^+ Sr^{2+} Ba^{2+} Tl^+	Group IA NH_4^+
SOLUBILITY <0.1 mol/L (SLIGHTLY SOLUBLE)	None	None	None	None	Ag^+	Ag^+ Pb^{2+} Hg_2^{2+} Cu^+ Tl^+	Ca^{2+} Sr^{2+} Ba^{2+} Ra^{2+} Ag^+ Pb^{2+}	Most	Most	Most

Periodic Table of Elements

KEY



- 🟥 elements do not form ionic compounds
- 🟩 gas
- 🟦 liquid
- ☐ solid

GROUP IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1 H 1.01 2.1 −259 1+ −253 1− 0.071 2 hydrogen																	2 He 4.00 − −272 −269 0.126 helium
3 Li 6.94 1.0 180 1+ 1347 0.53 lithium	4 Be 9.01 1.5 1278 2+ 2970 1.85 beryllium											5 B 10.81 2.0 2079 2550 2.34 boron	6 C 12.01 2.5 3550 4827 2.26 carbon	7 N 14.01 3.0 −210 3− −196 0.81 2 nitrogen	8 O 16.00 3.5 −218 2− −183 1.14 2 oxygen	9 F 19.00 4.0 −220 1− −188 1.505 2 fluorine	10 Ne 20.18 − −249 −246 1.20 neon
11 Na 22.99 0.9 97.8 1+ 883 0.97 sodium	12 Mg 24.31 1.2 649 2+ 1090 1.74 magnesium											13 Al 26.98 1.5 660 3+ 2467 2.70 aluminum	14 Si 28.09 1.8 1410 2355 2.33 silicon	15 P 30.97 2.1 44.1 w 3− 280 w 1.82 w 4 phosphorus	16 S 32.06 2.5 113 2− 445 2.07 8 sulfur	17 Cl 35.45 3.0 −101 1− −35 1.56 2 chlorine	18 Ar 39.95 − −189 −186 1.40 argon
19 K 39.10 0.8 63.7 1+ 774 0.86 potassium	20 Ca 40.08 1.0 839 2+ 1484 1.55 calcium	21 Sc 44.96 1.3 1541 3+ 2831 3.0 scandium	22 Ti 47.90 1.5 1660 4+ 3287 3+ 4.54 titanium	23 V 50.94 1.6 1890 5+ 3380 4+ 6.1 vanadium	24 Cr 52.00 1.6 1857 3+ 2672 2+ 7.19 chromium	25 Mn 54.94 1.5 1245 2+ 1962 4+ 7.43 manganese	26 Fe 55.85 1.8 1535 3+ 2750 2+ 7.87 iron	27 Co 58.93 1.8 1495 2+ 2870 3+ 8.9 cobalt	28 Ni 58.71 1.8 1453 2+ 2732 3+ 8.9 nickel	29 Cu 63.55 1.9 1083 2+ 2567 1+ 8.96 copper	30 Zn 65.38 1.6 420 2+ 907 7.13 zinc	31 Ga 69.74 1.6 29.8 3+ 2403 5.90 gallium	32 Ge 72.59 1.8 937 4+ 2830 5.32 germanium	33 As 74.92 2.0 − 3− 515 5.73 arsenic	34 Se 78.96 2.4 217 2− 685 4.79 selenium	35 Br 79.90 2.8 −7.2 1− 58 3.12 2 bromine	36 Kr 83.80 − −157 −152 2.41 krypton
37 Rb 85.47 0.8 38.9 1+ 688 1.53 rubidium	38 Sr 87.62 1.0 769 2+ 1384 2.5 strontium	39 Y 88.91 1.3 1522 3338 4.47 yttrium	40 Zr 91.22 1.4 1852 4377 6.51 zirconium	41 Nb 92.91 1.6 2468 4742 8.6 niobium	42 Mo 95.94 1.8 2617 4612 10.2 molybdenum	43 *Tc* 98.91 1.9 2172 4877 11.5 technetium	44 Ru 101.07 2.2 2310 3900 12.4 ruthenium	45 Rh 102.91 2.2 1966 3727 12.4 rhodium	46 Pd 106.40 2.2 1552 2+ 3140 4+ 12.0 palladium	47 Ag 107.87 1.9 962 1+ 2212 10.5 silver	48 Cd 112.41 1.7 321 2+ 765 8.65 cadmium	49 In 114.82 1.7 157 3+ 2080 7.31 indium	50 Sn 118.69 1.8 232 4+ 2270 2+ 7.31 tin	51 Sb 121.75 1.9 631 3+ 1750 5+ 6.70 antimony	52 Te 127.60 2.1 450 2− 990 6.24 tellurium	53 I 126.90 2.5 114 1− 184 4.93 2 iodine	54 Xe 131.30 − −111 −108 3.52 xenon
55 Cs 132.91 0.7 28.4 1+ 678 1.90 cesium	56 Ba 137.33 0.9 725 2+ 1640 3.5 barium	57 La 138.91 1.1 921 3457 6.15 lanthanum	72 Hf 178.49 1.3 2227 4602 13.3 hafnium	73 Ta 180.95 1.5 2996 5425 16.7 tantalum	74 W 183.85 1.7 3410 5660 19.3 tungsten	75 Re 186.21 1.9 3180 5627 21.0 rhenium	76 Os 190.20 2.2 3045 5027 22.6 osmium	77 Ir 192.22 2.2 2410 4130 22.5 iridium	78 Pt 195.09 2.2 1772 4+ 3827 2+ 21.5 platinum	79 Au 196.97 2.4 1064 3+ 2807 1+ 19.3 gold	80 Hg 200.59 1.9 −39.0 2+ 357 1+ 13.6 mercury	81 Tl 204.37 1.8 304 1+ 1457 3+ 11.85 thallium	82 Pb 207.19 1.8 328 2+ 1740 4+ 11.4 lead	83 Bi 208.98 1.9 271 3+ 1560 5+ 9.8 bismuth	84 Po (209) 2.0 254 2+ 962 4+ (9.2) polonium	85 At (210) 2.2 302 1− 373 − 2 astatine	86 Rn (222) − −71 −62 4.4 radon
87 Fr (223) 0.7 (27) 1+ (677) 2.4 francium	88 Ra 226.03 0.9 700 2+ 1140 5.0 radium	89 Ac (227) 1.1 1050 3+ 3200 10.07 actinium	104 (261) Rf rutherfordium Ku kurchatovium	105 (262) Ha hahnium Nl nielsbohrium	106 (263)												

58 Ce 140.12 1.1 799 3+ 3426 6.67 cerium	59 Pr 140.91 1.1 931 3+ 3512 6.77 praseodymium	60 Nd 144.24 1.2 1021 3+ 3068 7.00 neodymium	61 *Pm* (145) − 1080 3+ 2460 7.22 promethium	62 Sm 150.35 1.2 1077 3+ 1791 7.52 samarium	63 Eu 151.96 − 822 3+ 1597 2+ 5.24 europium	64 Gd 157.25 1.1 1313 3+ 3266 7.90 gadolinium	65 Tb 158.93 1.2 1356 3+ 3123 8.23 terbium	66 Dy 162.50 − 1412 3+ 2562 8.55 dysprosium	67 Ho 164.93 1.2 1474 3+ 2695 8.80 holmium	68 Er 167.26 1.2 159 3+ 2863 9.07 erbium	69 Tm 168.93 1.2 1545 3+ 1947 9.32 thulium	70 Yb 173.04 1.1 819 3+ 1194 6.97 ytterbium	71 Lu 174.97 1.2 1663 3+ 3395 9.84 lutetium
90 Th 232.04 1.3 1750 4+ 4790 11.7 thorium	91 Pa 231.04 1.5 < 1600 5+ − 15.4 protactinium	92 U 238.03 1.7 1132 6+ 3818 4+ 19.1 uranium	93 *Np* 237.05 1.3 640 5+ 3902 19.5 neptunium	94 *Pu* (244) 1.3 641 4+ 3232 6+ − plutonium	95 *Am* (243) 1.3 994 3+ 2607 13.7 americium	96 *Cm* (247) − 1340 3+ − − curium	97 *Bk* (247) − − 3+ − 14 berkelium	98 *Cf* (251) − − 3+ − − californium	99 *Es* (254) − − − − einsteinium	100 *Fm* (257) − − − − fermium	101 *Md* (256) − − − − mendelevium	102 *No* (259) − − − − nobelium	103 *Lr* (257) − − − − lawrencium

A LESSON RECORD FORM MUST BE COMPLETED FOR EVERY LESSON SUBMITTED FOR CORRECTION, AS ILLUSTRATED BELOW

A Lesson Record form with the **correct** label attached **must** be enclosed with **every lesson** submitted for correction, as illustrated below.

Correct use of these labels will ensure prompt processing and grading of your lessons.

The enclosed **Lesson Labels** must be checked for spelling and address details.

Please advise the Alberta Correspondence School promptly of any changes in name, address, school, or any other details and we will issue a revised set of labels. Your file number is permanently assigned and **must** be included on all correspondence with the Alberta Correspondence School. If the proper label and Lesson Record Form is not attached to each lesson as indicated it will delay your lessons being processed and credited to you.

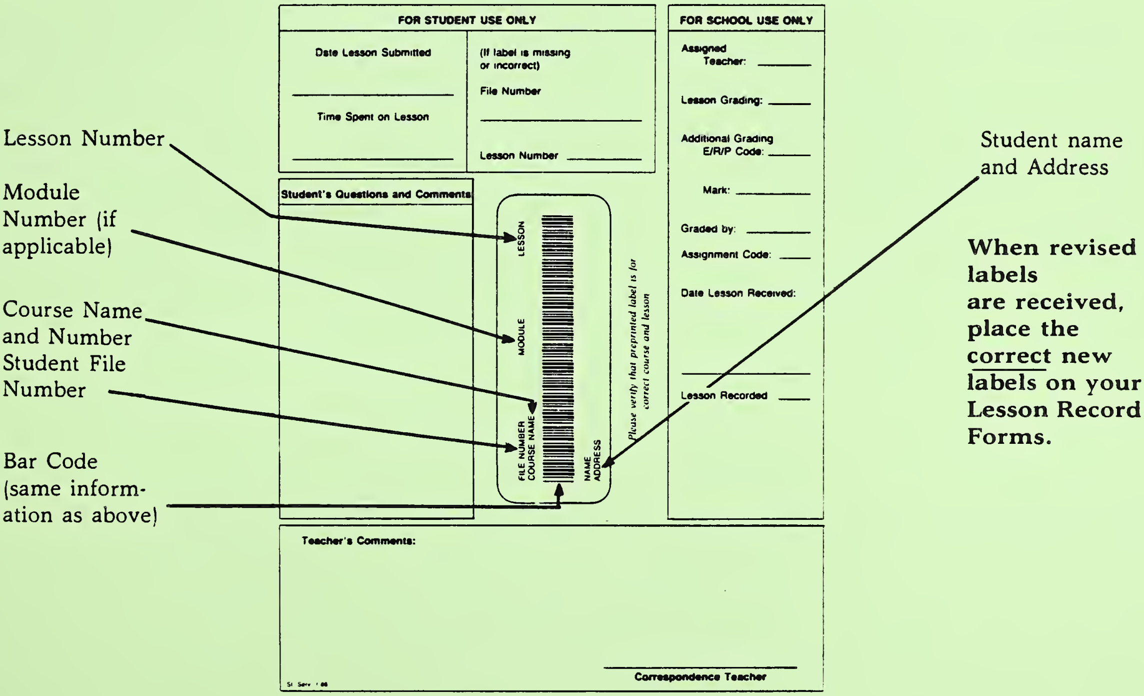
Lesson labels are to be attached to the **lesson record forms** in the space provided for student name and address.

Check carefully to ensure that the **subject name, module number** and **lesson number** on each label corresponds exactly with the lesson you are submitting.

Labels are to be **peeled** off waxed backing paper and **stuck** on the lesson record form.

Only **one** label is to be placed on each lesson.

LESSON RECORD FORM



FOR STUDENT USE ONLY

Date Lesson Submitted ______

Time Spent on Lesson ______

(If label is missing or incorrect)

File Number ______

Lesson Number ______

Student's Questions and Comments

Please verify that preprinted label is for correct course and lesson

FOR SCHOOL USE ONLY

Assigned Teacher: ______

Lesson Grading: ______

Additional Grading E/R/P Code: ______

Mark: ______

Graded by: ______

Assignment Code: ______

Date Lesson Received: ______

Lesson Recorded ______

Teacher's Comments:

Correspondence Teacher

St Serv 1 88

When revised labels are received, place the correct new labels on your Lesson Record Forms.

DO NOT MARK OR COVER BAR CODING.

CHANGE OF ADDRESS

If the address on your lesson record form differs from the address you supplied on your registration application, please explain. Indicate whether the different address is your home, school, temporary or permanent change of address.

LESSON RECORD FORM

1240 Chemistry 10

Revised 90/06

FOR STUDENT USE ONLY

Date Lesson Submitted

____

Time Spent on Lesson

____

(If label is missing
or incorrect)

File Number

____

Lesson Number __________

FOR SCHOOL USE ONLY

Assigned
Teacher: ________

Lesson Grading: ________

Additional Grading
E/R/P Code: ________

Mark: ________

Graded by: ________

Assignment Code: ______

Date Lesson Received:

____

Lesson Recorded ______

Student's Questions and Comments

Apply Lesson Label Here

Name

Address

Postal Code

Please verify that preprinted label is for correct course and lesson.

Teacher's Comments:

____

Correspondence Teacher

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ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

(1) All pages are numbered and in order, and no paper clips or staples are used.
(2) All exercises are completed. If not, explain why.
(3) Your work has been re-read to ensure accuracy in spelling and lesson details.
(4) The Lesson Record Form is filled out and the correct lesson label is attached.
(5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

First Class

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COMPLETE THIS QUESTIONNAIRE AND RETURN WITH LESSON ONE

Chemistry 10 Course Code 1240

Name ______________________________ Age ☐☐ File Number ☐☐☐☐☐☐☐

Address ______________________________ Phone Number ☐☐☐☐☐☐☐

______________________________ Postal Code ☐☐☐ ☐☐☐

1. Name of present school (if in attendance). ______________________________

2. Give your final mark (if any) in the following subjects:

Science 9 ____
Mathematics 9 ____
Chemistry 10 ____
Chemistry 20 ____
Chemistry 30 ____
Mathematics 10 ____
Mathematics 20 ____
Mathematics 30 ____
Biology 10 ____
Biology 20 ____
Biology 30 ____
Physics 10 ____
Physics 20 ____
Physics 30 ____

3. Do you plan to purchase a laboratory kit? __________

4. List any special qualifications or handicaps (jobs, illness, disabilities, etc.) which may influence your progress in this course. ______________________________

5. What is your reason for enrolling in this course? ______________________________

6. List any other subjects you are studying by correspondence instruction. ______________________________

Lesson	In	Out	Grade		Lesson	In	Out	Grade
1					A			
2					B			
3					C			
4					D			
5					E			
6					F			
7					G			
8					H			
9					I			
10					J			
11								
12								
13								
14								
15								
16								
17								
18								
19								
20								
EXAM								

Lesson Grade ☐

Exam Portion ☐

Lesson Portion ☐

Final Grade ☐

CONTENTS

A. Measurement of Matter

Measurement of matter helps us obtain quantitative information. In this lesson we will be interested in measuring length, mass, time and temperature.

B. Systems of Measurement

The system of measurement that used to be commonly used in Canada and is used in the United States now is the English system where the foot is the fundamental unit of length; the pound is the fundamental unit of mass; the second is the fundamental unit of time and Fahrenheit is the fundamental unit of temperature. This system is not very satisfactory for scientific work and has been replaced by the International System of Units used by scientists throughout the world.

In 1960, the International System of Units was established as a result of a long series of international discussions. The modernized metric system called SI, from the French name Le' Systeme International d' Unites, is now a general world trend. Canada has decided to convert to SI, which includes familiar metric units such as the metre, kilogram and Celsius temperature scale.

You probably have purchased your favorite brand of toothpaste in millilitre containers or cloth by the metre or gasoline by the litre. Farm land is now measured by the hectare instead of acres, and industry in general is converting to metric measurements.

Metric Units of Length

1 millimetre (mm) = 0.001 metre (m)
1 centimetre (cm) = 0.01 metre (m)
1 decimetre (dm) = 0.1 metre (m)
1 kilometre (km) = 1000 metre (m)

Note: In this section and in the ones on metric units of volume and mass, the letters enclosed in brackets are the abbreviations for the units.

Metric Units of Area

The common units are the square centimetre (cm^2) and the square metre (m^2). For example, a floor 4 metres long and 2 metres wide has an area of $4\ m \times 2\ m = 8\ m^2$.

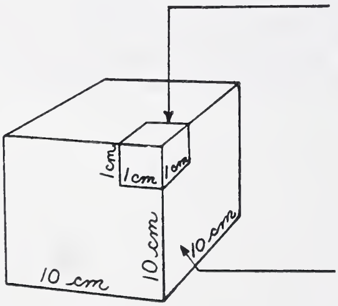
Metric Units of Volume

The common units are the cubic centimetre (cm^3), the cubic metre (m^3), the litre (L), and the millilitre (mL). For example, a box 3 centimetres long, 2 centimetres wide, and 2 centimetres deep has a volume of $3\ cm \times 2\ cm \times 2\ cm = 12\ cm^3$. The litre is the volume of a cubic box 10 centimetres on each edge. That is, 1 litre $= 10\ cm \times 10\ cm \times 10\ cm = 1000\ cm^3$ and 1/1000 part of the litre equals 1 millilitre which in turn is equal to $1\ cm^3$.

Note: Avoid using the text's (ℓ) and (ml) for litre and millilitre.

Metric Units of Mass

When the metric system was devised the units of length, volume, and mass were chosen so as to be directly related to one another. The unit of mass, the kilogram (kg), was defined as the mass of $1000\ cm^3$ or of one litre of pure water at 4°C. This is equivalent to defining the density $\left(\frac{\text{mass}}{\text{volume}}\right)$ of water at 4°C to be 1 gram per cm^3 $\left(\frac{g}{cm^3}\right)$ or 1 kilogram per litre $\left(\frac{kg}{L}\right)$. For example, a box $2\ cm \times 3\ cm \times 6\ cm$ would have a volume of $36\ cm^3$ or 36 millilitres and would hold 36 grams of water at 4°C.

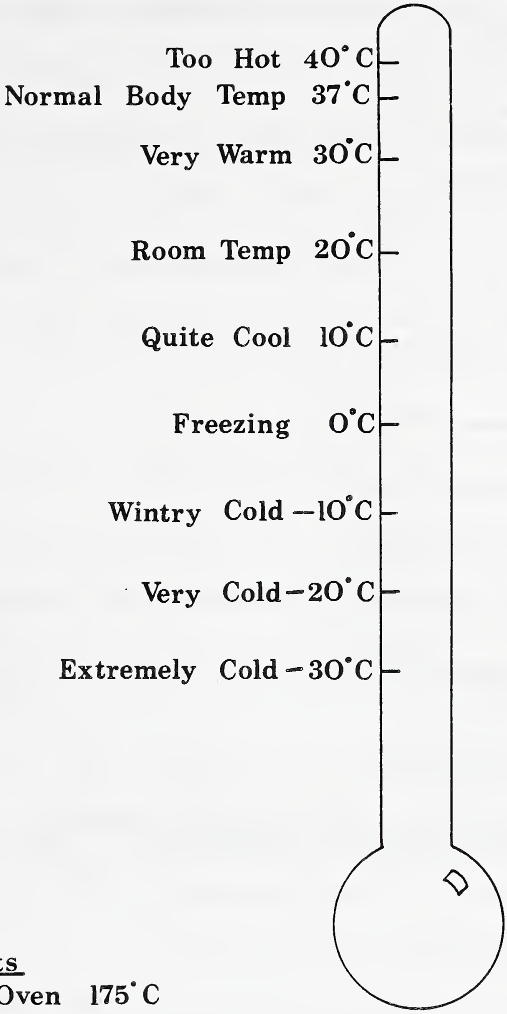


$1\ cm^3$ or 1 mL contains 1 gram (1 g) of water at 4°C.

1 gram has 1000 milligrams (1000 mg).

$1000\ cm^3$ or 1 litre contains 1 kg of water at 4°C.

Common Temperature Units



Other Units

Moderate Oven	**175°C**
Water Boils	**100°C**

To do: Please Read Sections 1-6 and 1-7 (pages 12-16) of your textbook.

Do Practice Exercise on Page 30 - Problem 1

Self Test on Page 33 - Problem 5

Check answers with textbook. DO NOT send for correction.

You **must** send in for correction all exercises and experiments in the lessons. In general, if the answers are provided you do not send them in for correction. Otherwise the exercises must be sent in.

Exercise 1 (To be sent for correction)

Using the information given in this lesson and Appendix 1, page 285, give the metric equivalents for the following:

(a) 454 g = ________________ kg

(b) 50 g = ________________ mg

(c) 3 L = ________________ cm^3

(d) 16 mL = ________________ L

(e) 15 mm = ________________ cm

(f) 150 cm^3 = ________________ L

(g) 2.5 kg = ________________ g

(h) 4.5 L = ________________ mL

(i) 60 km = ________________ m

(j) 76 m = ________________ cm

Exercise 2 (To be sent for correction)

1. The dimensions of a "Black Box," which will be used later in an experiment, were recorded as follows: length 16 cm, width 10 cm, height 80 mm.

   Calculate the volume of this box.

2. (a) A dollar bill is about ____________ cm × ____________ cm.

   (b) The average weight of a new born baby is about: (underline)

   (i) 300 kg, (ii) 30 kg, (iii) 3 kg

   (c) The thickness of a dime would be about

   (i) 5 mm, (ii) 1 mm, (iii) 0.1 mm

C. The Concept of Matter and Energy

To do: Read sections 1-3, 1-4, and 1-5 (pages 9 to 11 of your textbook) and then complete the following exercise.

Exercise 3 (To be sent in for correction)

1. Matter exists in three forms (a) ____________________

   (b) ____________________ (c) ____________________

2. Compare the characteristics of the three forms of matter by filling in the table.

Form		Characteristics
(a) Solid	(i)	______________________________
	(ii)	______________________________
	(iii)	______________________________
		______________________________
(b) Liquid	(i)	______________________________
	(ii)	______________________________
	(iii)	______________________________
		______________________________
(c) Gas	(i)	______________________________
	(ii)	______________________________
	(iii)	______________________________

3. What is meant by the Law of Conservation of Energy?

4. Albert Einstein showed the relationship between matter and energy. Write the equation that shows this relationship.

5. Your textbook gives several samples of how energy can be changed from one form to another. Give two additional examples below.

   (a)

   (b)

D. Significant Figures

To do: Read Appendix 3 (pages 287-290) of your textbook. This section will enable you to determine the number of significant figures and how to use them in addition, subtraction, multiplication and division of numbers.

To do: Practice Exercise - page 30 - problems 2, 3, 4 and 5

Self Test - page 33 - problem 6

Do not send for correction.
Check answers in textbook.

Exercise 4 (To be sent for correction)

Complete the following statements:

Significant figures include all the digits which are certain and one more about which there is ______________________. Final zeros to the left of an understood decimal point ______________________ (are or are not) significant. Final zeros to the left of a decimal point ______________________ (are or are not) significant. All zeros between non-zero digits ______________________ (are or are not) significant. The number 4807 has __________ significant digits. Final zeros to the right of a decimal point ______________________ (are or are not) significant.

How many significant figures are there in each of the following?

1. (a) 48.1 __________
(b) 1206 __________
(c) 26.0 __________
(d) 0.0001 __________
(e) 0.1255 __________
(f) 36.090 __________
(g) 80.00 __________
(h) 3050 __________
(i) 350. __________
(j) 52 000. __________

2. Addition and Subtraction

Give the answers to the correct number of significant figures in each case, rounding off where necessary.

(a)
$$\begin{array}{r} 1.28 \\ 0.49 \\ \underline{0.062} \end{array}$$

(b)
$$\begin{array}{r} 0.001 \\ 0.082 \\ \underline{0.03} \end{array}$$

(c)
$$\begin{array}{r} 6\ 400 \\ 375 \\ \underline{25\ 950} \end{array}$$

(d)
$$\begin{array}{r} 95.3 \\ \underline{-\ 6.15} \end{array}$$

(e)
$$\begin{array}{r} 3.006 \\ -\ \underline{1.4} \end{array}$$

(f)
$$\begin{array}{r} 43.43 \\ \underline{-\ 0.0029} \end{array}$$

3. Multiplication and Division

(Use correct number of significant figures)

(a) 12.9 × 1.23 = __________
(b) 2.72 × 3.1 = __________
(c) 71.45 ÷ 0.99 = __________
(d) 9 ÷ 21.4 = __________
(e) 12 × 15.86 = __________

E. Exponential Numbers

To do: Read Appendix 2 of the textbook (page 286) and all the extra material on exponential numbers found in this lesson.

After studying the examples carefully, complete the exercises on page 15 of this lesson.

When dealing with either very large numbers (such as 489 000 000 000) or very small numbers (such as 0.000 000 000 32) it is more convenient to express them as exponential numbers.

e.g. $489\ 000\ 000\ 000 = 4.89 \times 10^{11}$

$0.000\ 000\ 000\ 32 = 3.2 \times 10^{-10}$

Numbers written in exponential form are also said to be written using scientific notation.

Study the table below.

$1000 = 10^{3}$ ←-------- exponent is 3

$100 = 10^{2}$ ←-------- exponent is 2

$10 = 10^{1}$

$1 = 10^{0}$

$\frac{1}{10} = 0.1 = 10^{-1}$

$\frac{1}{100} = 0.01 = 10^{-2}$

$\frac{1}{1000} = 0.001 = 10^{-3}$

$0.000\ 000\ 1 = 10^{-7}$

Note:

$$\boxed{\begin{aligned} 10^{2} &= \frac{1}{10^{-2}} \\ 10^{-2} &= \frac{1}{10^{2}} \end{aligned}}$$

Any number may be expressed as an integral power of ten or as the product of two numbers one of which is an integral power of ten.

e.g. $34\ 500 = 3.45 \times 10^{4}$

Rule 1: A shift of the decimal point to the left requires the use of a positive exponent.

e.g. $1985 = 198.5 \times 10^{1}$

$= 19.85 \times 10^{2}$

$= 1.985 \times 10^{3}$

Rule 2: A shift of the decimal point to the right requires the use of a negative exponent.

e.g.

$$1.985 = 19.85 \times \frac{1}{10} = 19.85 \times 10^{-1}$$

$$1.985 = 198.5 \times \frac{1}{10^2} = 198.5 \times 10^{-2}$$

$$0.000\ 000\ 48 = \frac{4.8}{10\ 000\ 000} = \frac{4.8}{10^7} = 4.8 \times 10^{-7}$$

Many measurements in science involve very large or very small numbers which often need to be multiplied or divided. So that this can be done conveniently, such numbers should be expressed as two factors. The first factor known as the <u>digital factor</u> is a number between 1 and 10 or a number with one integer to the left of the decimal point e.g. 2.45, 3.6, 1.0 and 8.007. The other factor is called the <u>exponential factor</u> and is the correct power of 10.

<u>Examples</u>

NUMBER	=	Digital Factor	×	Exponential Factor
185	=	1.85	×	10^2
1850	=	1.85	×	10^3
185 000	=	1.85	×	10^5
0.185	=	1.85	×	10^{-1}
0.000 185	=	1.85	×	10^{-4}

Notice that the digital factor in all these numbers is the same. Only the exponential factor varies.

The reason for converting numbers to the exponential form is that the numbers between 1 and 10 are by far the easiest to visualize and to work with. The difficulty due to the largeness or smallness of a number is removed by the use of exponents.

<u>Example</u> Multiply 20 000 by 0.04.

While this can be done by the usual multiplication, it must be admitted that there is a chance of misplacing the decimal point.

Using exponents :

$$20\ 000 = 2 \times 10^4$$

and $$0.04 = 4 \times 10^{-2}$$

$$20\ 000 \times .04 = 2 \times 10^4 \times 4 \times 10^{-2}$$

$$= 2 \times 4 \times 10^{4+(-2)} \quad \text{(Add exponents)}$$

$$= 8 \times 10^2$$

$$= 800$$

Exercise 5 (To be sent in for correction)

1. Change the following to non-exponential form.

(a) 3×10^2 = 300

(b) 8.6×10^4 = __________

(c) $1.867\ 32 \times 10^5$ = __________

(d) 1.71×10^{-2} = 0.0171

(e) 9.4×10^{-6} = __________

(f) 8.01×10^1 = __________

(g) 0.015×10^{-1} = __________

(h) 20×10^3 = __________

(i) 46.5×10^{-3} = __________

(j) 0.05×10^2 = __________

2. Change the following digital factor to correspond to the appropriate power of ten. Follow the 3 examples done for you.

(a) 448×10^4 = 4.48 $\times 10^6$

(b) 0.386×10^6 = 0.0386 $\times 10^7$

(c) 332×10^{-4} = 3.32 $\times 10^{-2}$

(d) 324.5×10^9 = __________ $\times 10^{11}$

(e) 8.1×10^6 = __________ $\times 10^4$

(f) 224×10^{-5} = __________ $\times 10^{-3}$

(g) 2.24×10^3 = __________ $\times 10^2$

(h) 5.55×10^{-2} = __________ $\times 10^{-1}$

(i) 1.0×10^5 = __________ $\times 10^6$

(j) 6.02×10^{23} = __________ $\times 10^{24}$

3. Express each of the following in standard exponential form; that is, as the product of two numbers, one of which is an exponential factor of ten, the other number being a digital factor (having one integer to the left of the decimal). Several examples have been done for you.

(a) 25 000 = $\underline{2.5 \times 10^4}$

(b) 495 = ________

(c) 0.015 = ________

(d) 0.495 = ________

(e) 0.005 = ________

(f) 0.000 094 = $\underline{9.4 \times 10^{-5}}$

(g) 0.000 78 ________

(h) 80.10 = ________

(i) 2658 = ________

(j) 201 = ________

Addition and Subtraction of Exponential Numbers

Rule: Convert all numbers to the same power of ten before carrying out addition or subtraction.

Example: Add 7.62×10^3 cm and 5.2×10^2 cm

$5.2 \times 10^2 \text{ cm} = 0.52 \times 10^3 \text{ cm}$

Therefore

$$\begin{array}{r} 7.62 \times 10^3 \text{ cm} \\ \underline{0.52 \times 10^3 \text{ cm}} \\ 8.14 \times 10^3 \text{ cm} \end{array}$$

Note: Regardless of what power of ten you work with, your answer should be expressed as the product of the digital factor (a number between 1 and 10) and ten to the appropriate power.

Exercise 6 (To be sent in for correction)

Complete the following problems, using the correct number of significant figures in all cases.

1. A student was given the mass of three samples of compound. Sample A: 2.5×10^2 g, Sample B: 0.53×10^3 g and Sample C: 300.0 g. Calculate the total mass of the three samples of compound using exponential numbers.

2. An evacuated glass bulb weighed 14.1×10^1 g. Upon filling it with a gas at room temperature and pressure it weighed 1.55×10^2 g. Calculate the mass of the gas in the bulb, using exponential numbers.

3. Add or subtract the following exponential numbers. Express your answer in standard exponential form as explained previously. (Round off to the correct number of significant digits)

(a) $2.3 \times 10^4 + 4.6 \times 10^4 =$ ____________________

(b) $2.5 \times 10^4 + 5.3 \times 10^5 =$ ____________________

(c) $10.8 \times 10^{12} - 4.6 \times 10^{12} =$ ____________________

(d) $8.6 \times 10^7 - 3.0 \times 10^6 =$ ____________________

(e) $2.56 \times 10^{-2} + 0.36 \times 10^{-3} =$ ____________________

(f) $3.6 \times 10^2 - 2.6 \times 10^{-1} =$ ____________________

(g) $6.02 \times 10^{23} + 0.301 \times 10^{24} =$ ____________________

(h) $4.2 \times 10^2 + 5.1 \times 10^1 + 6.5 \times 10^3 =$ ____________________

Multiplication and Division of Exponential Numbers

Rule: In multiplication, exponents of like bases are added.

Examples

(a) $x^2 \times x^5 = x^{2+5} = x^7$

(b) $10^2 \times 10^3 = 10^{2+3} = 10^5$

(c) $10^7 \times 10^{-2} = 10^{7+(-2)} = 10^5$

(d) $(3 \times 10^2) \times (5 \times 10^3) = (3 \times 5)(10^{2+3}) = 15 \times 10^5$
$= 1.5 \times 10^6$

(e) $(2 \times 10^3) \times (4 \times 10^{-2}) = (2 \times 4) \left[10^{3+(-2)}\right] = 8 \times 10^1$

(f) $2000 \times 3500 = (2 \times 10^3)(3.5 \times 10^3)$
$= (2 \times 3.5)(10^{3+3}) = 7 \times 10^6$

(g) $x^2 \times 10^3$ = (can't be multiplied)

Rule: In dividing, exponents of like bases are subtracted.

Examples

(a) $\frac{x^3}{x^2} = x^{3-2} = x^1 = x$

(b) $\frac{10^4}{10^2} = 10^{4-2} = 10^2$

(c) $\frac{10^5}{10^7} = 10^{5-7} = 10^{-2} = \frac{1}{10^2}$

(d) $\frac{8 \times 10^3}{2 \times 10^{-5}} = \frac{8}{2} \times 10^{3-(-5)} = 4 \times 10^{3+5} = 4 \times 10^8$

(e) $\frac{3.6 \times 10^{-2}}{1.6 \times 10^3} = \frac{3.6}{1.6} \times 10^{-2-3} = 2.3 \times 10^{-5}$

(f) $\frac{x^2}{10^3}$ = (can't be divided)

<u>Note</u>: An expression with an exponent of zero is equal to one.

$10^0 = 1, \qquad a^0 = 1, \qquad 8.2 \times 10^0 = 8.2$

Why?

$$x^a \times \frac{1}{x}a = x^{a-a} = x^0$$

But

$$x^a \times \frac{1}{x}a = 1$$

$$\therefore x^0 = 1$$

Exercise 7 (To be sent in for correction)

1. Study examples on the previous page. Multiply and express your answer in standard exponential form, using the correct number of significant figures.

(a) $3.6 \times 10^2 \times 4.0 \times 10^3 = 14.4 \times 10^{2+3} = 14.4 \times 10^5 =$ ____________

(b) $8.2 \times 10^{-3} \times 2.2 \times 10^5 = 18.04 \times 10^{-3+5} =$ ____________

(c) $10^3 \times 10^6 =$ ____________

(d) $3 \times 10^7 \times 2 \times 10^4 =$ ____________

(e) $1.2 \times 10^{12} \times 7.0 \times 10^8 =$ ____________

(f) $2.5 \times 10^8 \times 4.0 \times 10^{-3} =$ ____________

(g) $1.3 \times 10^{-6} \times 3.0 \times 10^{-4} =$ ____________

2. Divide and express your answer in standard exponential form.

(a) $10^8 \div 10^2 =$ ____________

(b) $2.7 \times 10^8 \div 3 \times 10^5 =$ ____________

(c) $4.5 \times 10^{-9} \div 5 \times 10^4 =$ ____________

(d) $15.6 \times 10^{-8} \div 1.3 \times 10^{-12} =$ ____________

(e) $6.52 \times 10^{23} \div 14 \times 10^{-3} =$ ____________

END OF LESSON 1

LESSON RECORD FORM
1240 Chemistry 10
Revised 90/06

FOR STUDENT USE ONLY

Date Lesson Submitted ____________________

Time Spent on Lesson ____________________

(If label is missing
or incorrect)

File Number ____________________

Lesson Number __________

FOR SCHOOL USE ONLY

Assigned
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Lesson Grading: __________

Additional Grading
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Lesson Recorded ______

Student's Questions and Comments

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Please verify that preprinted label is for correct course and lesson.

Teacher's Comments:

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Correspondence Teacher

St. Serv. 21-89

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CONTENTS

A. Introduction to Laboratory Work

The emphasis of this lesson is upon experimentation. You are the scientists. Your observations on the behaviour of matter should grow into conclusions which are exciting. The laboratory investigations you will carry on in this course are planned to (1) acquaint you with common laboratory equipment and apparatus, (2) familiarize you with experimental techniques and procedures, (3) provide data which can serve as basis for discussion.

(1) Preparing for Laboratory Work

It is very important that you understand the directions and purpose of each experiment before starting to work on it. It is a good idea to read the procedure through once before starting the experiment. As you read, decide what data you will have to record and prepare data tables. Then when you start the experiment, your observations can be recorded systematically, which makes the record meaningful for future use. Your record of experiments should be a complete but brief outline of the experiment. It is not necessary to copy out the instructions. What you did and what you saw are noted under the heading: "Record". The more precise your work is, the more valuable it will be.

In recording, use the passive voice and be impersonal.

Recommended	Not Recommended
The test tube was heated.	I heated the test tube. Heat the test tube.

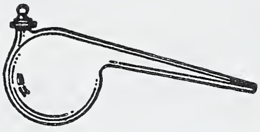
(2) Apparatus - see list of common laboratory apparatus enclosed on page 2 of this lesson.

(3) Study Rules of Conduct and Safety in the laboratory in your Laboratory Manual (pages 3 to 4).

These rules will be emphasized again as you proceed with your experiments.


BUNSEN BURNER


ERLENMEYER FLASK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BEAKER


MORTAR AND PESTLE


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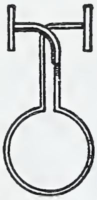
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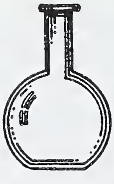

FLORENCE FLASK


DEFLAGRATING SPOON


TONGS

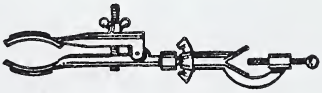

TEST TUBE BRUSH


THISTLE TUBE


GRADUATED CYLINDER


CRUCIBLE AND COVER

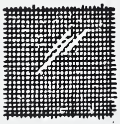

EVAPORATING DISH

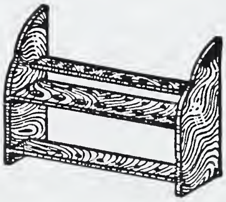

BURETTE CLAMP


FUNNEL


CALCIUM CHLORIDE TUBE


WIDE MOUTHED GAS BOTTLE

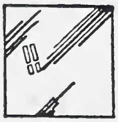

WIRE GAUZE


TEST TUBE RACK

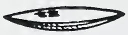

IGNITION TUBE


TEST TUBE


TEST TUBE HOLDER


GLASS PLATE


TRIANGLE


WATCH GLASS


BLOW PIPE

Exercise 1

The following exercise is based on information found in your Laboratory Manual, pages 5 to 19, under the heading: Laboratory Techniques and Procedures.

Fill in the blanks with the appropriate term or phrase.

1. After lighting a gas burner adjust the burner so that the flame is

   ______________________________.

2. The ____________________________ cone of the flame is the hottest part of the flame.

3. When weighing chemicals on a balance the metal weighing pan should be protected by ________________________________.

4. List three types of apparatus used in measuring the volume of liquids.

   (a) _________________________________

   (b) _________________________________

   (c) _________________________________

5. The surface of liquids when viewed in a glass cylinder is always curved. This lens-shaped surface is called a(n) ___________________________.

6. For a most accurate reading of a liquid in a graduated cylinder the meniscus must be viewed along ____________________________________

   ____

7. In transferring solids from reagent bottles list three general rules to be observed in order to avoid contamination of reagents.

   (a) ____

   (b) ____

   (c) ____

   ____

8. In transferring liquids from reagent bottles, remove the stopper and

____

9. In filtering solids from their liquid material the liquid that passes through the filter paper is called a(n) ________________________

____

10. Why is it important to wash the residue after filtering?

____

____

11. A(n) ________________ may be used to separate a precipitate from a liquid in a short period of time.

12. When a solid dissolves in a liquid (salt dissolved in water) the solid can be recovered by ____________________________.

13. What is meant by "fire-polishing" of glassware?

____

____

____

14. List two practical purposes of fire-polishing.

(a) ______________________________________

(b) ______________________________________

B. Experiment 1 QUALITATIVE AND QUANTITATIVE OBSERVATION

To do: Read Section 1-2 (pages 3 to 9) of your textbook, then do the experiment.

In this experiment you will get the opportunity to make good observations which are the key to success in science.

Observations may be of two types:

1. Qualitative
2. Quantitative

Description and examples of these two types of observations appear in your Laboratory Manual, page 21. Please read this material carefully and proceed with the experiment.

Follow the outline below:

Note: The experiment has been modified to some extent. We hope this change will provide more opportunity for scientific observation.

Purpose: (state briefly)

____

____

Materials:

4 sugar cubes
4 glasses or paper cups
hot and cold tap water
watch or clock

Procedure:

Please read through the procedure before attempting the experiment.

Place one sugar cube into a glass or paper cup as directed below:

Glass 1 - 3/4 full of cold water

Glass 2 - 3/4 full of hot water

Glass 3 - 1/4 full of cold water

Glass 4 - use pulverized (crushed) sugar cube in 3/4 full of cold water

Set up your apparatus in advance. Place sugar cubes in rapid succession in order that more accurate comparisons can be made in a given interval of time.

Exercise 2

Tabulate your observations in the table below.

Time:	Cold Water	Hot Water	Limited Water	Crushed-Cold
20 s (seconds)				
2 min				
5 min				
10 min				
25 min				
40 min				
60 min				

Exercise 3

1. From your observations, please write the word Qualitative or Quantitative after each of the statements given below.

(a) In hot water sugar dissolves faster than in cold water.

____

(b) In 10 min exactly one-half of the sugar cube melts.

____

(c) All cubes started dissolving after a short period of time.

____

(d) A sugar cube, before dissolving, is 1.5 cm × 1.5 cm.

____

(e) Bubbles form as sugar cubes dissolve. ____________________________

(f) In 50 min the sugar cube in hot water dissolves completely.

____

2. Fill in the blanks with appropriate terms or sentences.

(a) Did you observe any bubbles rising from the whole sugar cubes?

____

(b) Did you observe any bubbles rising from the crushed sugar cube?

____

(c) Give a possible interpretation for the occurrence of bubbles when sugar cubes dissolve.

____

____

____

C. <u>Graphs</u>

To do: (i) Please read Sections 1-9 and 1-10 (pages 20 to 24) of your textbook, on how to draw and interpret graphs.

(ii) Do Practice Exercise on page 31 - problem 8 (a-e).

Check your answers with the textbook after you have completed the exercise.

D. Experiment 2 MELTING AND FREEZING OF SUBSTANCES

Prelaboratory Study for Experiment 4

Consider the heating and cooling curves in the time-temperature relationship for ice shown in a sample experiment below. Also read the procedure for the lab.

Record:

Ice from a refrigerator was heated over a limited range by a constant source of heat. The temperature of the sample confined in a test tube was noted every one-half minute.

Similarly water in a test tube was cooled and frozen and the temperature noted every one-half minute.

Data:

Phase Change: ice → water

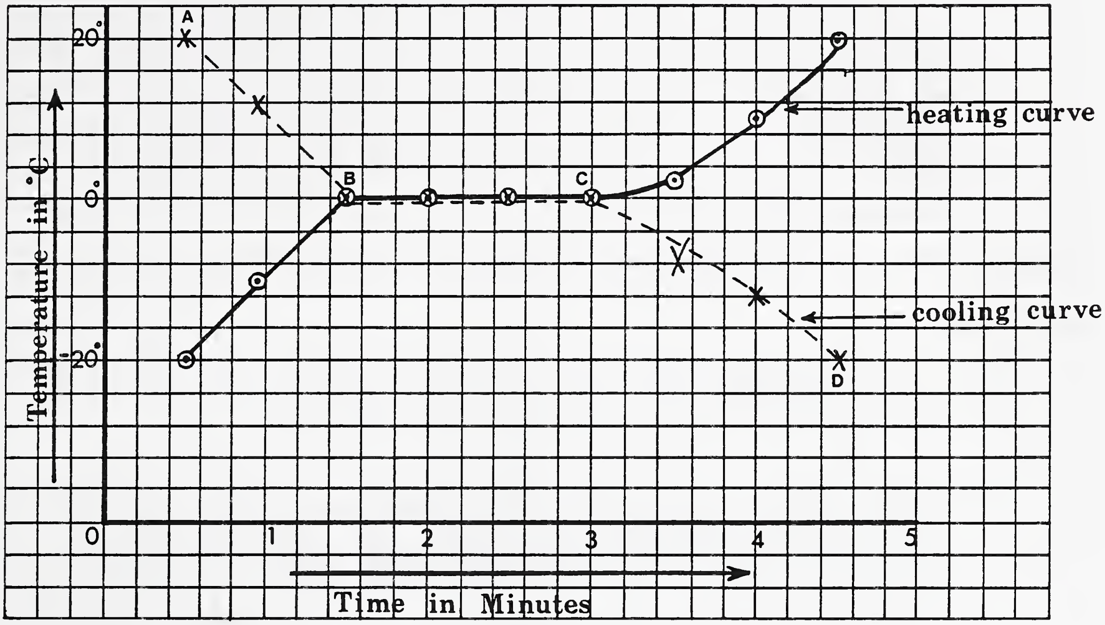
Time Minutes	Heating data °C	Phase
0.5	-20°	solid
1.0	-10°	solid
1.5	0°	phase change
2.0	0°	
2.5	0°	
3.0	0°	
3.5	2°	liquid
4.0	10°	liquid
4.5	20°	liquid

Phase Change: water → ice

Time Minutes	Cooling data °C	Phase
0.5	20°	liquid
1.0	12°	liquid
1.5	0°	phase change
2.0	0°	
2.5	0°	
3.0	0°	
3.5	-8°	solid
4.0	-11°	solid
4.5	-20°	solid

From the data a graph was obtained by plotting the time against the temperature.

The cooling and heating curves obtained were:



Interpretation of curves:

AB on the cooling curve shows the time taken for the water to cool from 20°C to 0°C. At 0°C, the freezing point, ice starts to appear. With further cooling, freezing continues until all the water is frozen. While freezing is taking place, the temperature remains constant as shown by BC, the plateau on the graph. Continued cooling causes a drop in temperature of the ice as shown by CD on the graph.

A similar curve is shown for the heating of ice. Note that there is also a plateau, or a line of constant temperature, on the heating curve. This represents the time required for all the ice at 0°C to change to water at 0°C.

Increasing the amount of ice used in the experiment would not change the melting temperature but the slope of the heating curve would be more gradual and the plateau longer. Applying more heat per minute would cause a more rapid rise in temperature and would shorten the length of the plateau.

Observe that the melting and freezing temperatures for water are the same. This is typical for a pure substance. For mixtures, melting and freezing often occur over a range of temperatures so that a plateau will not be seen on the graph.

Read the procedure for the experiment.

Experiment 2

Materials

* paradichlorobenzene
* test tube
* test tube holder
thermometer
* beaker
water
saucepan (pot)

Object: To study the temperature changes as pure substances melt and freeze.

Procedure:

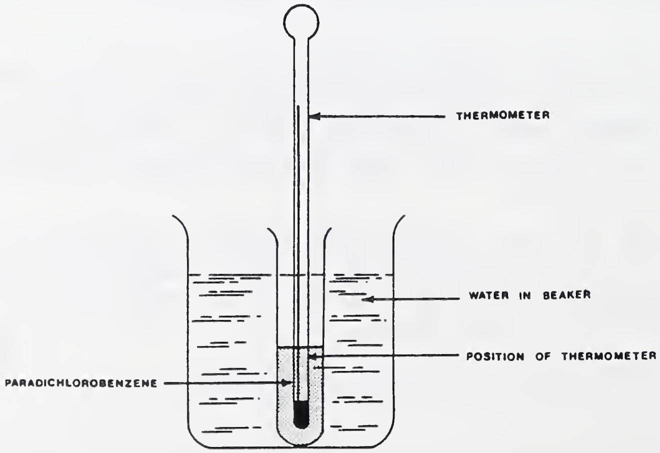
1. (a) Fill a saucepan (pot) three quarters full of water. Heat until boiling, remove from heat, and turn off heat.

(b) Fill the beaker three quarters full of lukewarm tap water.

(c) Place the test tube of paradichlorobenzene into the saucepan. When the paradichlorobenzene has melted insert the thermometer into the test tube. When the temperature reaches about 70° C place the test tube and contents into the beaker of lukewarm water. On a piece of looseleaf labelled **Cooling** record the temperatures at thirty second (30 s) intervals and the phase until the temperature drops to about 40°C. Do **NOT** remove the thermometer from the solid paradichlorobenzene (the thermometer may break and it is required to be stuck in the solid for the next part of this lab).

2. (a) Reheat the water in the saucepan until the water starts to steam but not actually boil. Turn the heat to its lowest setting so that the water will stay hot.

(b) Place the test tube containing the solid paradichlorobenzene and thermometer in the hot water. (Be careful not to get water into the test tube.) On a piece of looseleaf labelled **Heating** record the temperatures at 30 s intervals and the phase until the temperature reaches about 65°C.



Your data should have a time interval during which the temperature hardly changed similar to the data found on page 8. **If your data does not have this "plateau" then you should repeat the experiment.** Maybe the heating or cooling was too rapid and you will have to make modifications to the amount of water or temperature of water accordingly.

When you have reasonable results, copy the data neatly onto the table provided. Indicate solid (s), liquid (l), or both (b), under phase.

DATA TABLE

Cooling			Heating		
Time (Minutes)	Temperature (°C)	Phase	Time (Minutes)	Temperature (°C)	Phase
0.0			0.0		
0.5			0.5		
1.0			1.0		
1.5			1.5		
2.0			2.0		
2.5			2.5		
3.0			3.0		
3.5			3.5		
4.0			4.0		
4.5			4.5		
5.0			5.0		
5.5			5.5		
6.0			6.0		
6.5			6.5		
7.0			7.0		
7.5			7.5		
8.0			8.0		
8.5			8.5		
9.0			9.0		
9.5			9.5		
10.0			10.0		

Organizing the Information Collected:

The graph paper provided in this lesson will be used to record your data in graphical form. Follow sample graph on page 9 of this lesson and suggestions given below. Indicate time along the horizontal axis (the abscissa) using five spaces for one minute of time. Indicate temperature on the vertical axis (the ordinate) allowing five spaces for each ten degrees.

Plot heating and cooling temperatures on the same graph, starting your plotting from the left edge of the graph. Use a tiny circle for each heating temperature and a tiny (x) for each cooling temperature. Join the points for the cooling curve with a blue line; those for the heating curve with a red line.

Please be sure the graph has a title as well as labels on the axis.

Interpretations:

Exercise 4

1. Explain why there is a plateau on your graph.

____

____

____

2. Based on your data:

   (a) what is the melting point of paradichlorobenzene? ____________

   (b) what is the freezing point of paradichlorobenzene? ____________

3. What effect would increasing the amount of paradichlorobenzene have on the shape of the heating and cooling curves on your graph?

____

____

____

4. Explain how you could show a material is a mixture and not a pure substance?

____

____

____

Conclusion:

A mixture will have a range of freezing and melting points. A pure substance will have one melting and freezing point. You can conclude that paradichlorobenzine is a pure substance since it has one distinct melting and freezing point.

Exercise 5

Which of the following are observations and which are interpretations.

(a) The plateau on the cooling curve of a liquid indicates a phase change. ______________________________

(b) Sugar dissolves faster in hot water. ______________________________

(c) Bubbles rise from dissolving sugar cubes.

____

(d) A precipitate is formed when $Pb(NO_3)_2$ is mixed.

____

(e) The precipitate formed is lead iodide.

____

(f) Heat and light are produced by a burning candle.

____

(g) The bubbles that rise from the dissolving sugar cubes are air trapped in the cubes, and not gas due to reaction, since no bubbles formed with crushed sugar.

____

E. Estimating Volumes of Liquids

(Exp. 1-2, Page 23 - Laboratory Manual)

Exercise 6

1. Number of drops per cm^3

You will need an eye dropper, water, and a 10 cm^3 graduated cylinder, (your graduated cylinder will have markings in mL's. This is the same as cm^3 or cubic centimeter.) Using your eye-dropper count the number of drops it takes to make 1 cm^3 of water in a 10 cm^3 graduated cylinder. Try it several times, and give the average in the space provided.

There are ____________ drops in 1 cm^3 of water.

2. Estimating the Volume in a Test Tube

(a) Hold a 13 mm × 100 mm test tube at eye level. Place the lower part of your thumb on the test tube at the level you think would indicate 1 cm^3 of water. (b) Pour water into the test tube to the level marked by your thumb.

(c) Pour this water into a 10 cm^3 graduated cylinder. See how close your estimate came to being 1 cm^3.

(d) Repeat this process for 2 cm^3 and 5 cm^3.

(e) After you have practiced your estimates above, estimate the total number of cm^3 in the test tube when completely filled. Give your estimate in the blank below.

Estimated ____________ cm^3 in a completely filled test tube.

(f) Using the result in 1 calculate the total number of drops of water in a completely filled test tube.

(g) Finally, using your graduated cylinder, measure out the exact number of cm^3 of water needed to completely fill the test tube. How does this result compare with your estimated result in 2(e)?

F. Tests for Some Common Gases

Exercise 7

Read the information given in your Laboratory Manual - Part II (Pages 24 and 25) and briefly describe the test for the following gases:

1. Hydrogen:

2. Oxygen:

3. Carbon dioxide:

Note: You may have experienced some of these tests in former science courses.

END OF LESSON 2

LESSON RECORD FORM

1240 Chemistry 10

Revised 90/06

FOR STUDENT USE ONLY

Date Lesson Submitted ______________________

Time Spent on Lesson ______________________

(If label is missing
or incorrect)

File Number ______________________

Lesson Number ____________

FOR SCHOOL USE ONLY

Assigned
Teacher: ________

Lesson Grading: ________

Additional Grading
E/R/P Code: ________

Mark: ____________

Graded by: ____________

Assignment Code: ______

Date Lesson Received: ______________________

Lesson Recorded ______

Student's Questions and Comments

Apply Lesson Label Here

Name ______________________

Address ______________________

____

____

Postal Code ______________________

Please verify that preprinted label is for correct course and lesson.

Teacher's Comments:

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Correspondence Teacher

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ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

(1) All pages are numbered and in order, and no paper clips or staples are used.
(2) All exercises are completed. If not, explain why.
(3) Your work has been re-read to ensure accuracy in spelling and lesson details.
(4) The Lesson Record Form is filled out and the correct lesson label is attached.
(5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

First Class

Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope. Correspondence lessons will travel faster if first-class postage is used.**

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

CLASSIFICATION OF MATTER

A. Physical and Chemical Properties
B. Physical and Chemical Changes
C. Experiment 3: Obtaining Hydrogen using Magnesium
D. Classification of Matter
  (1) homogeneous or heterogeneous
  (2) metals or nonmetals
  (3) mixtures or pure substances
  (4) ionic or molecular substances
E. Density
F. Experiment 4: Density
G. Experiment 5: Flow Charts

A. Physical and Chemical Properties

To do: Read section 2-1 in the textbook.

To summarize the above section, if you take any material and do anything with it, without involving any other chemical, you are measuring or determining physical properties. Density (or heaviness), hardness, melting point and other such qualities are examples of physical properties since no other chemicals are involved.

On the other hand, when describing chemical properties, you must involve at least two substances and indicate how they are capable of reacting with one another. For instance, the gases at the far right hand side of the periodic table are inert which means that they are not capable of reacting with other substances. Substances which are capable of reacting with other substances are active. How readily one substance interacts with another substance to produce a third substance is a measure of its chemical properties.

To do: Practice Exercise on page 56 - problem one. Then check your answers on page 309. (These self-checks do not have to be sent in for correction.)

If you had many errors, perhaps you should read the above lesson notes and section 2-1 again. When you feel that you understand physical and chemical properties, do the exercises below.

Exercise 1

Classify the following properties as either physical or chemical (P or C).

(a) Water boils at 100°C _____

(b) Oxygen readily reacts with other substances. _____

(c) Iodine can dissolve in carbon tetrachloride. ______

(d) Sodium can react with water to produce hydrogen. ______

(e) Gold can be pounded into extremely thin sheets. ______

(f) Silver tarnishes easily at room temperature. ______

B. Physical and Chemical Changes

To do: Read Section 2-2 in the textbook.

When a piece of paper is cut in half, this is a physical change since no new material is formed. The above section talks about phase changes. A phase change occurs when a solid melts, when a liquid evaporates, when a vapour condenses or when a liquid freezes. At room temperature, as you probably know, oxygen and hydrogen are gases while water is a liquid. Water (H_2O) is composed of hydrogen and oxygen. What happens when water evaporates from a dish? (a) Do the water molecules turn into hydrogen and oxygen gas, or (b) do the water molecules remain as water molecules, but in effect dissolved in air (just like sugar dissolves in tea)? (a) or (b)? ________

You probably know that oxygen makes everything burn brighter and faster which is the reason that no smoking is allowed in hospital rooms where oxygen is being used, whereas water dampens fires. Would something burn brighter or not so brightly if it were very foggy outside? What would you call a phase change: a chemical change or a physical change? ________________

When ice is heated, water results. When water is cooled, ice results. When white bread is put into a toaster, toast results, but what happens when toast is put into a fridge? Do you get white bread back? Obviously, a chemical change occurred when bread was toasted but only a physical change occurred when ice melted.

A chemical change occurs when one or more substances react to produce a new substance which has entirely different chemical and physical properties. A physical change occurs when a single substance is deformed in some way.

To do: Practice Exercise on Page 56 - Problem 2 and then check your answers. If you feel you understand this material, do the exercise below.

Exercise 2

Classify the following changes as physical or chemical changes. (P or C)

(a) boiling of water at 100°C _____

(b) evaporation of water at 20°C _____

(c) dissolving sugar in water _____

(d) exploding dynamite _____

(e) completely blending several pieces of fruits and vegetables in a blender _____

(f) the digestion of the completely blended food _____

(g) the rusting of iron _____

(h) the burning of gasoline _____

(i) a sample of ocean water is evaporated and a white solid residue is obtained _____

(j) milk, on a hot day, turns sour _____

(k) hydrogen and oxygen are put into one container but no different substances are produced _____

(l) hydrogen and oxygen are ignited by a spark to produce water vapour _____

(m) heating dough to make a cake

hint: will cooling the cake afterwards produce the dough again? _____

C. Experiment 3 OBTAINING HYDROGEN USING MAGNESIUM

To do: Read Exp. 2-2 in the Laboratory Manual, but do not follow any procedures or worry about any questions on this experiment since we will do things a bit differently.

Due to postal regulations, we cannot send any acid or base in our Lab Kit which is stronger than a 0.1 M solution, so whether you are using a Lab Kit or a school lab, please follow the instructions below.

Exercise 3

Using page 39 of the Laboratory Manual as a guide, record six physical properties of magnesium ribbon from the lab kit (or use a small piece (≈4 cm) of magnesium ribbon from your own lab).

1. ______________________________
2. ______________________________
3. ______________________________
4. ______________________________
5. ______________________________
6. ______________________________

Taking some household vinegar, record its odor, color, relative density compared with water (ie. is it much less, about the same, or much greater?)

1. odor: ______________________________
2. colour: ______________________________
3. density: ______________________________

Materials:

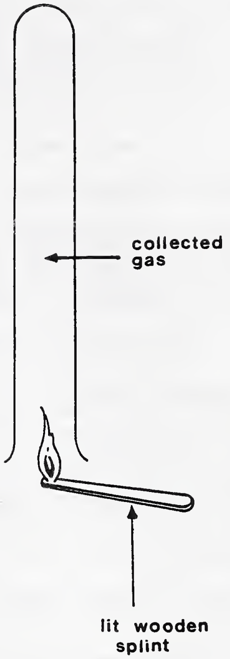
* two 13 × 100 mm test tubes
* a small piece of Mg ribbon (≈ 4 cm)
about a quarter of a test tube of vinegar
* a wooden splint
a match
* a beaker

Observe all safety precautions and be extra careful about pointing the test tube away from yourself if you don't wear glasses or have safety glasses.

Procedure:

1. Fold the magnesium ribbon and put it in the bottom of one test tube.
2. Put about a quarter of a 13 × 100 mm test tube of vinegar in with the magnesium ribbon.

3. Immediately catch all bubbles with the other test tube inverted over the first test tube. (Also check if the bottom of the test tube which contains vinegar and magnesium feels any warmer.)
4. When the bubbling has stopped put the test tube containing the vinegar into a beaker so the contents do not spill out.
5. Light a wooden splint and hold it underneath the top test tube which should still be inverted. (Hopefully, you heard a popping noise after a second or two.)
6. After putting the splint out, pour the contents of the other test tube into a clean beaker or glass jar, and let all water evaporate (you may want to speed it up by putting the clear solution over a heat register.)
7. Check it for tiny crystals after all of the water has evaporated.



Exercise 4

1. List at least three things which indicate that a chemical change has taken place.

   (a) ______________________________

   (b) ______________________________

   (c) ______________________________

2. The popping noise resulting from the wooden splint indicates the presence of hydrogen. Where did this hydrogen come from? (You started off with magnesium and vinegar which is a combination of acetic acid and water.)

   ____

3. If you had too much magnesium ribbon and too little acetic acid, some of the magnesium ribbon would not have reacted. True or False __________

4. Which type of change caused the test tube to get warmer, chemical or physical? (Underline the correct answer.)

4. (b) If the test tube got warm, the reaction was exothermic. If the test tube became colder, the reaction was endothermic. Which type of reaction did you have?

____

5. What happened to the magnesium ribbon during the reaction?
   (a) It escaped as a gas with the bubbling.
   (b) It evaporated with the water.
   (c) It formed a salt and was in the tiny crystals after the water had evaporated. ______

D. Classification of Matter

(1) homogeneous or heterogeneous
(2) metals or nonmetals
(3) mixtures or pure substances
(4) ionic or molecular

To do: Read sections 2-3 and 2-5 in the textbook.

Homogeneous substances look alike throughout. For example, a mixture of gases in a bottle or a completely dissolved component in a liquid results in a homogeneous solution. When you add sugar to tea and do not stir it, the sugar will settle to the bottom. This would be a heterogeneous solution. When the sugar is completely stirred, the solution is homogeneous.

The textbook lists several physical properties of metals but it also notes that some physical properties of metals and nonmetals may be the same. For example mercury, a metal, and bromine, a nonmetal, are both liquids at room temperature. You might ask, what makes a metal a metal? The distinguishing feature of all metals is that they all have similar chemical properties, even though physical properties may greatly differ. All metals readily lose their outer electrons in chemical reactions whereas nonmetals gain electrons in chemical reactions. (Electrons will be covered in greater detail later.)

Mixtures and pure substances are different in that a mixture will always consist of several pure substances mixed together, either homogeneously or heterogeneously. Water is a pure substance as well as sugar but when the two are mixed a mixture forms even though you may not be able to tell the difference between pure water and water with sugar dissolved in it.

To discuss the difference between ionic and molecular substances may be premature since electrons have not been discussed yet. For now, all you need to know is that electrons are negative charges and protons are positive charges. If a substance has an equal number of electrons and protons it is neutral and is therefore a molecular substance. An ionic substance has either lost electrons or it has gained electrons. If a substance, such as a metal, loses electrons, it will have more protons than electrons and it will be positively charged. If a nonmetal gains electrons, resulting in more electrons than protons, you will have an ion with a negative charge.

Exercise 4: Classification of Matter

1. Classify the following as either mixtures or pure substances: sugar, salt, a watermelon, a shovel full of dirt, the contents of a pencil sharpener, vinegar, distilled water, mercury, milk, and a flask filled with air.

Mixtures	Pure Substances

2. Classify the following as either homogeneous or heterogeneous: (A) sugar and water well stirred, (B) well stirred equal volumes of oil and water, (C) well stirred mixture of mercury and water, (D) a mixture of ice and water, (E) vinegar, (F) a well mixed mixture of marbles and salt, (G) a 0.0001 mol/m^3 solution of HCl (in lab kit), (H) distilled water, (I) a shovel full of dirt. (Use apprpriate letters.)

homogeneous	heterogeneous

3. List five substances, or parts of substances in your home, which are metallic.

(a) ______________________________

(b) ______________________________

(c) ______________________________

(d) ______________________________

(e) ______________________________

4. List five substances, or parts of substances in your home, which are non metallic.

(a) ______________________________

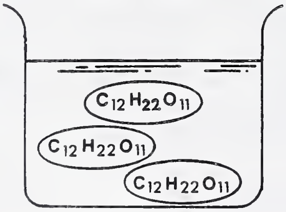
(b) ______________________________

(c) ______________________________

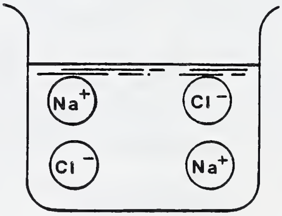
(d) ______________________________

(e) ______________________________

5. The following diagrams indicate how sugar ($C_{12}H_{22}O_{11}$) and how salt (NaCl) dissolves in water.



sugar solution



salt solution

(a) Which is an ionic solution? ______________________________

(b) Which is a molecular solution? ______________________________

E. Density

Apply the formula: $\text{Density} = \frac{\text{mass}}{\text{volume}}$ in your problem solving.

To do: Read Section 1-11 (pages 24 - 26) of your textbook.

Practice Exercise on Page 31 - Problem 6. Do not send for correction. Check your answers at the back (page 307).

Exercise 5

Solve the following problems. (Densities given on pages 25 and 26 are to be used.)

1. Calculate the mass of 20 cm^3 of copper.

2. Calculate the volume of 28.3 g of lead.

3. A piece of metal, 2 cm × 3 cm × 1.5 cm weighs 24.3 g. Calculate its density and state what metal it is.

F. Experiment 4

DENSITY

To do: Read Exp. 1 4 - Page 31 of the Laboratory Manual

Materials:

10 cm^3 graduated cylinder (lent)
balance (lent to students)
iron nail (5 to 7 cm long)
2 brass masses from your balance set (5 and 2 g masses)
* aluminum rod

Purpose:

To provide practical experience in finding the density of objects.

Procedure:

Obtain the four objects described above. Before doing the experiment predict which object has the greatest density and which has the lowest.

Density	Object
Highest	
Lowest	

Using your balance find the mass of each object. Record it in the table below.

	Object	Mass
1.	iron nail	
2.	5 g mass	
3.	2 g mass	
4.	aluminum rod	

To find the volume of the nail use the following procedure. Fill the 10 cm^3 cylinder with water to a height slightly more than the length of your nail. Take a careful measure of this height and record. (Use table provided). Next, carefully place the nail into the cylinder. The nail should be totally immersed in water. Measure and record the new height of water in the cylinder. (See page 10 of Laboratory Manual.) The difference between these two heights will give you the volume of the nail. (The nail displaces its own volume.)

To obtain the volume of each other object proceed in a similar manner and record your data in the table below.

Object*	Height of Water Before Immersion	Height of Water After Immersion	Volume of Object (Amount of Water Displaced)
iron nail			
5 g mass			
2 g mass			
aluminum rod			

*If you are not using one of our lab kits, indicate your replacements above.

Exercise 6

Analysis of Data

Using your data, answer the questions which follow. If you find it necessary, review the section on density.

1. Calculate the density of the iron nail, or your substitute object.

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

2. How does your experimental result compare with the given value for iron (or your object) on page 26 of your textbook?

____

____

3. What possible reason could you give for the difference in values?

____

____

4. Calculate the density of the 5 g mass (or your second object).

5. Calculate the density of the 2 g mass (or your third object).

6. Should the results obtained in question (4) and (5) be similar? Explain.

____

7. Calculate the density of the aluminum rod (or fourth object).

8. How does this compare with the expected value?

____

____

9. How did your predictions for highest and lowest density (page 10) compare with experimental results?

____

____

____

____

H. Experiment 5 (To be sent for correction)

FLOW CHARTS

In this experiment you will learn how to use a flow chart. The experiment also introduces some important laboratory procedures, techniques, and some terms you need to know. The actual understanding of all chemical reactions is not important at this time; that understanding will come later.

Materials:

* lead nitrate ($Pb(NO_3)_2$)
water
* sodium iodide (NaI)
10 cm^3 graduated cylinder
* 13 × 100 mm test tube
* 2 pieces of filter paper
* 2 beakers
balance (lent)

Procedure:

Using your balance weigh 0.5 g of lead nitrate, $Pb(NO_3)_2$.
Caution: Do not place chemicals on the metal weighing pan without first protecting the metal with a piece of plastic or lightweight paper. Many chemicals will corrode metals and cause deterioration of the pan. (See other hints, page 8 - Laboratory Manual).

Place the weighed compound into a dry, clean beaker.

Measure out 25 cm^3 of water using your graduated cylinder. (See page 10 - Laboratory Manual under "Meniscus".) Pour water into the beaker to dissolve the compound. Stir thoroughly. Note the color. Weigh 0.2 g of sodium iodide, NaI, and place it in a clean beaker. Add 10 cm^3 of water. Stir until compound dissolves. Pour this solution into the beaker containing the previously dissolved compound $Pb(NO_3)_2$. Note the color of the

precipitate. ____________________

Obtain a clean glass or paper cup and proceed with filtration. Follow the method on page 13 - Laboratory Manual under "Filtering Precipitates" and note the common terms.

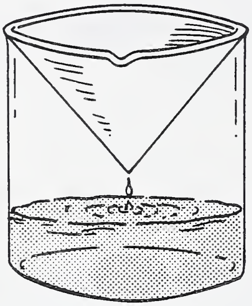
Since a funnel is not available, rest the filter paper against the side of a cup or glass. Be careful not to puncture the moistened filter paper. (See the diagram below)

Filtering Procedures:

1. Fold the filter paper in half.

2. Fold the paper once again so that it is quartered.

3. Since you don't have a funnel, place the filter paper in the beaker so it rests against the mouth of the beaker. See diagram.



You may have to hold the filter paper with your hand if it tends to slip inside the beaker.

4. Moisten the paper with water after placing it into the beaker to prevent it from popping out.

Note the color of the filtrate. If it is yellow, filter it again until the filtrate is clear. After filtration is complete, place the filter paper with residue in it to dry. What is the color of the residue in the filter paper?

____

Pour the filtrate into a glass beaker and gently heat the solution (use the stove or a candle) until evaporation is complete. What is the color of the residue that remains?

____

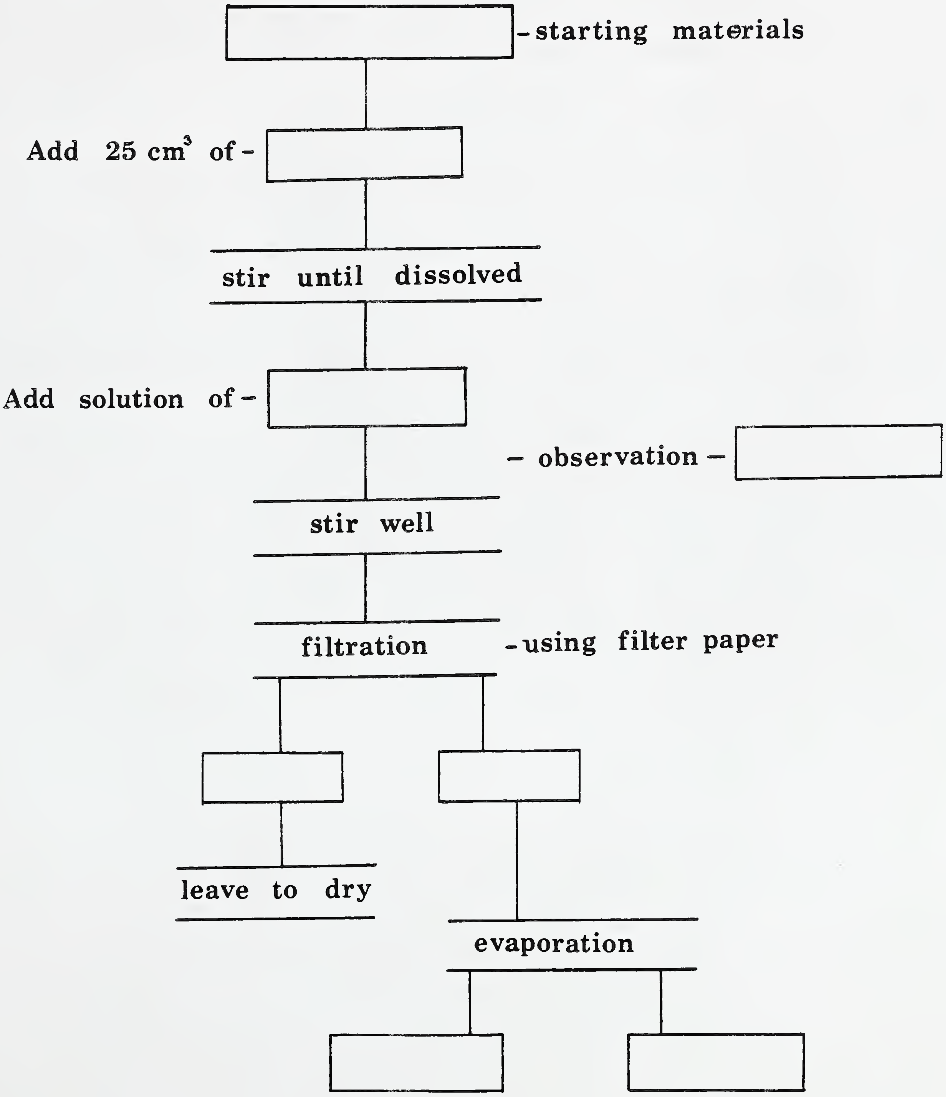
CAUTION: TREAT ALL CHEMICALS AND REAGENTS AS POISONOUS AND KEEP AWAY FROM REACH OF SMALL CHILDREN.

THE REMAINING LEAD NITRATE AND SODIUM IODIDE SHOULD BE STORED IN A SAFE PLACE FOR LATER USE IN EXPERIMENT 10.

Exercise 7

The whole operation of the above experiment can be represented in a flow chart which summarizes your procedures in a compact, easy-to-follow manner. To give you some practice in making up a flow chart, fill in the blocked spaces provided.

Flow Chart



If you encounter problems with this chart, please refer to the flow chart on page 26 - Laboratory Manual.

1. Given that $NaNO_3$ is soluble in water, and based on what you observed when trying to dissolve the compounds in this experiment, which of the following compounds is a precipitate

   (a) sodium iodide (NaI)
   (b) lead nitrate ($Pb(NO_3)_2$)
   (c) lead iodide (PbI_2)
   (d) sodium nitrate ($NaNO_3$) ________

2. What did evaporation accomplish? (This is the last step in the flow chart.)

   ____

   ____

   ____

3. Fill in the blanks with an appropriate term.

   (a) The solid caught by the filter paper is called a precipitate or

   ______________________.

   (b) Filtering helps to separate a ____________ from a

   ______________ material.

   (c) The liquid called a(n) ______________ passes through the filter paper.

END OF LESSON 3

LESSON RECORD FORM

1240 Chemistry 10

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CLASSIFICATION OF MATTER (continued)

E. Mixtures, Pure Substances, Elements and Compounds

To do: Review sections 2-3 and 2-5 in the textbook, and read section 2-4 in the textbook.

Although this section overlaps the previous section, the classification of matter will be dealt with a bit further. It was seen in the previous section that a pure substance has the same single substance throughout. You have 100% of one substance and nothing else in a pure substance. In a mixture one has at least two different substances present.

Pure substances can be further divided into elements and compounds. A compound contains two or more elements chemically combined in a definite proportion by weight. A compound can be broken down into its elements. When a compound is broken down into its elements, what kind of a change would this be, a physical change or a chemical change? ____________________ An element is a substance which cannot be broken down into simpler substances by chemical means.

When looking at pure substances, it is not easy to tell if you have an element or a compound without doing chemical tests. Since only two elements are liquid at room temperature, bromine (red) and mercury (silver), it is a safe bet that any other pure liquid you will come across at room temperature is a compound. Also, unless you have something made of a pure metal, such as copper wires or aluminum foil, it is safe to assume it is an alloy of two or more metals or a compound of some sort. All organic substances such as plastics, paper, wood, etc. are compounds of carbon and hydrogen with possibly oxygen and other elements. The name of the substance might also give it away as to whether something is an element or a compound. For instance, the coating on a tin can is "tin" which is one of the elements, but the coating on a fence is paint and "paint" is not listed as one of the elements. Of course, you will always find exceptions. Can you think of a form of pure carbon which is called by some other very familiar name? ____________________

To do: Page 56 - problems 3, 4, 5, 6. Page 57 - problems 1, 2, 3, 4, 5. Check answers in the back of the textbook.

Exercise 1

1. Divide the following pure substances into elements and compounds: water, gasoline, magnesium ribbon, paper, pure liquid oxygen, pure liquid acetic acid, zinc, glass, carbon monoxide, nickel, and hydrogen gas.

elements	compounds

2. Define an element.

3. Define a compound.

4. Since uranium is radioactive and is capable of converting itself to lead, would you classify uranium as an element or a compound?

Why?

5. How could you show that a salt water solution is a mixture and not a compound?

____

____

____

____

____

6. Mixtures can be separated into compounds or elements by simple physical or mechanical means. True or False ____________

7. Compounds can be separated into elements by more complex physical or mechanical means. True or False ____________

8. Pure acetic acid has 40% carbon, 6.7% hydrogen and 53.3% oxygen. If you had a substance which contains 30% carbon, 9% hydrogen and 61% oxygen, could it still be pure acetic acid? ____________ Why?

____

____

9. Milk has 87% water with minute amounts of minerals, fats, proteins and other substances. Would you still have milk if you had 92% water and 8% of the other substances? ____________ Why? ______________

____

____

10. Different elements have greatly different properties. It is important to use each element for its beneficial properties, but to be careful that any given element does not harm us. For instance, tungsten has an extremely high melting point, and when alloyed with steel, gives the steel great strength at high temperatures. Chromium, which is very corrosion resistant, is added to iron to produce stainless steel. Other metals, such as mercury and lead, have many good uses, but can also be harmful or poisonous

to man. Which metal is commonly used for wiring in a house? ____________ What is a very light pot likely to be made of? ____________ Which metal is often the filament of a light bulb? ____________ A galvanized pail is made of iron, and has a thin coating of ____________. Which element is put into balloons to make them float? ____________ Which element has to be put into a tank to enable sea divers to breath? ____________ Which element would be an important source of nuclear energy in power plants? ___uranium___ Seventy percent of the sun is composed of ___hydrogen___ .

11. State the Law of Definite Proportions. ____________

____

____

12. If a compound has 1 g of hydrogen to 16 g of oxygen, can you assume it is water? ____________ Why? ____________

____

F. Experiment 6 - FORCES BETWEEN MAGNETIC POLES

To do: Read Exp. 2-5, page 45

This lab was meant to demonstrate that forces between poles of a magnet increase as the poles come closer together. Instead of doing the experiment as it is listed, we will do a simplified version, but which should still get the same point across.

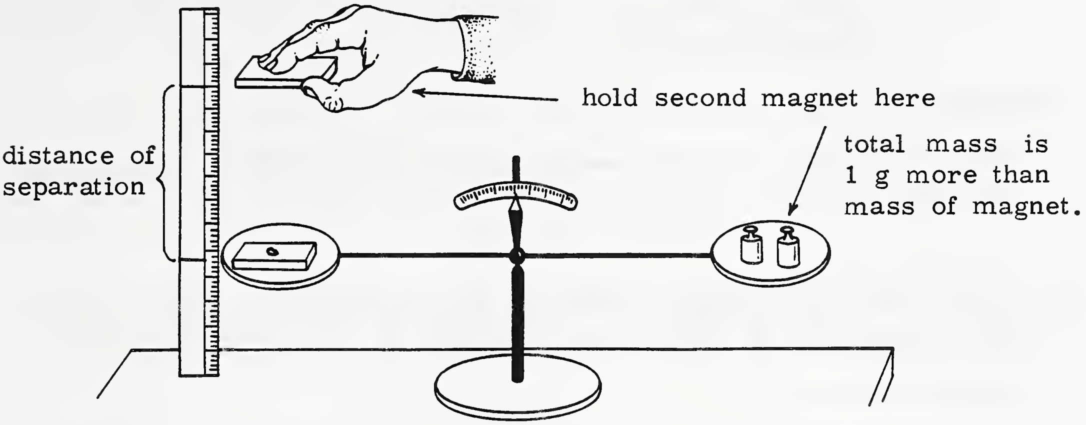
Materials:

* 2 ceramic magnets
ruler
balance

Procedure:

Lay one ceramic magnet on one side of the balance, and put enough masses on the other side to balance the mass of the magnet. Then add another gram to the side opposite the magnet. To do this, you may have to exchange some masses like taking a 1 g mass off and putting a 2 g mass on.

Then hold the second magnet above the first magnet in such a way that the magnets repel each other. Record the distances the 2 magnets are separated when the scale is balanced. (See diagram below.)



When recording the distance, you should go from the center of one magnet to the center of the next. After you have found the distance for 1 g, do the same for 2, 3, 4 etc. grams of excess weight on the right side.

excess weight	distance
1 g	
2 g	
3 g	
4 g	

To do the above experiment, using repulsion of magnets is easier than using attractive forces, since in this way, the magnets don't have to be clamped down. Of course you realize that attractive forces also become greater as the distance between the magnets is lessened.

G. Energy and Chemical Change

To do: Read Section 2-6

As you will learn in greater detail later, all elements are composed of negatively charged electrons and positively charged protons. Analogous to magnets, like charges repel, and unlike charges attract. In every compound, there is an optimum distance for the various elements to be separated, since various attractive and repulsive forces are at work. Basically, a bond forms because electrons of one element are attracted to the protons of the next element. Due to repulsive forces between two different nuclei, and electrons from two elements, the elements in a compound can't get too close together. Since both attractive and repulsive forces are acting in a compound, the situation is more complex than that of two magnets.

To do: Practice Exercise #11 and Self Test #7 on page 57, and then check your answers.

Exercise 2

1. Which has more potential energy: (a) an object which is 20 m above the ground, or (b) an object of the same mass which is 10 m above the ground?

   ____

2. Which has more potential energy: (a) two individual oxygen atoms or (b) a molecule of oxygen which consists of two atoms of oxygen?

   ____

3. When a chemical reaction occurs, the products may have less potential energy than the reactants. Since total energy is not lost in a chemical change, would you expect the products to be (a) hotter or (b) colder than the reactants?

   ____

4. When the distance between magnets is halved , the force of repulsion between the magnets will be:

   (a) twice as much
   (b) three times as much
   (c) four times as much ________

H. Symbols

To do: Read 2-8 in the textbook.

As you can see inside the front and back covers of this textbook, every element can be represented by a symbol which consists of one or two letters. There is no need to memorize these as such, but you will automatically learn the symbols of the most common elements by using them.

To do: Practice Exercise questions 7, 8, 10 - page 57 and Self Test questions 8 and 9 - page 58. Check your answers.

Exercise 3

Give the symbols of the nine most abundant elements in the earth's lithosphere and atmosphere.

_____________ _____________ _____________ _____________ _____________

_____________ _____________ _____________ _____________

I. Experiment 7 PRODUCTION OF FeS

To do: Read Exp. 2-4 in the Laboratory Manual.

If you wish to buy muriatic acid (impure hydrochloric acid) from the drug store or use acid from the school, you may do so, but the minimum that we require anyone to do is what can be done with equipment in the Lab Kit.

Materials:

* beaker
* test tube clamp
balance
* five 13 × 100 mm test tubes
* 0.1 molar HCl*
candle
matches
* sulfur
* iron filings
* magnet
hot and cold water
hammer
old rag

Procedure: Part I

Examine the sulfur and iron filings, and give five physical properties of each. (See pages 38 and 39 for physical properties.)

Exercise 4

Give 5 physical properties of sulfur.

1. ____________________ 4. ____________________
2. ____________________ 5. ____________________
3. ____________________

*As far as HCl is concerned, 0.1 molar = 0.1 normal. 1 M = 1 mol/L Other books using SI may use mole per cubic metre (mol/m^3). 1 mol/L = 0.001 mol/m^3.

Give 5 physical properties of iron filings.

1. ______________________________

2. ______________________________

3. ______________________________

4. ______________________________

5. ______________________________

Add 0.1 molar HCl to a small sample of sulfur in one test tube and add the same acid to a small sample of iron filings in another test tube. The reaction will be weak due to the dilute acid and hopefully you will see tiny bubbles form on the iron filings but no reaction on the sulfur. It may help to pour the acid into a test tube, and then put the test tube into a beaker of very hot water for a minute or two. The hotter acid would show more of a reaction. On the basis of the reaction with a metal an acid in Experiment 5, what do you suppose the gas on the iron filings is? ____________ Smell the gas. (It should be odorless)

Part II:

Mix 3.5 g of iron with 2 g of sulfur. List five physical properties of this mixture. (As one of the physical properties, compare the density of the mixture to the density of an equal volume of iron filings or an equal volume of sulfur.)

Physical properties of the mixture

1. ______________________________

2. ______________________________

3. ______________________________

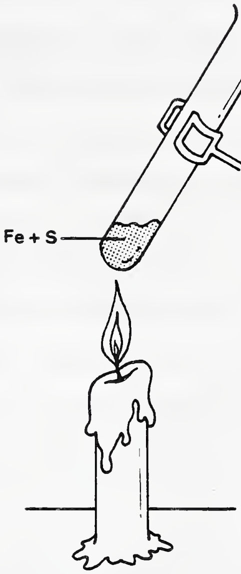
4. ______________________________

5. ______________________________

Put a small portion of this mixture into a test tube and add a bit of warmed up 0.1 molar HCl. Carefully smell the gas that results. What is the odor of the gas? ____________

Part III:

Grind up the mixture of iron filings and sulfur so that no lumps of sulfur remain, then add a portion (about a centimetre or less) into a test tube. Heat the contents of this test tube over a candle flame for a few minutes. (You could secure the candle by letting some wax from the burning candle fall on your working surface and then putting the base of the candle in this liquid wax.) Unfortunately, you won't be able to see a light in the test tube since a carbon coating from the burning candle obscures everything. This is nothing to worry about. After a few minutes, quench the test tube in a beaker of cold water and examine the contents of the test tube. You may have to break the test tuve to get the contents out. Be careful not to cut yourself. Observe the safety precautions on page 44 of the Lab Manual.



List five physical properties of this new compound.

1. ____________________

2. ____________________

3. ____________________

4. ____________________

5. ____________________

Do you suppose this new compound would have different chemical properties than the mixture of iron and sulfur? ______________. To find out, put this new compound in a test tube and add warmed up 0.1 molar HCl. Even when using the weak 0.1 molar solution HCl, you should be able to smell a very foul smelling odor. When using a stronger solution of HCl, be careful. The resulting gas (hydrogen sulfide) is extremely poisonous. Be sure to do the experiment by an open window or out of doors if possible. Should you spill any HCl on you, wash it off immediately with lots of water and apply baking soda.

If you got a foul smelling hydrogen sulfide (H_2S) gas from the warm dilute hydrochloric acid (HCl) (dissolved in water) and iron sulfide (FeS), where do you think the hydrogen came from? ______________________.
Where did the sulfur come from? ______________________.

Give two physical properties of the mixture of (Fe + S) which were the same as that of the compound FeS.

1. ______________________

2. ______________________

Give two physical properties of the mixture which were different from that of FeS.

1. ______________________

2. ______________________

Do you think the compound had different chemical properties than the mixture? __________ Give a reason for your answer. ______________

____

If you had mixed all of your 3.5 g of iron filings (Fe) with your 2 g of sulfur (S) you would have obtained 5.5 g of iron sulfide (FeS). What would have happened if you mixed and heated 3.5 g of Fe with 3.5 g of S?

(a) You would get 7 g of FeS.
(b) You would get 5.5 g FeS + 1.5 g S.
(c) No reaction would occur. ______________________

In your answer above, were you applying the Law of Definite Proportions?

__________.

From page 26 of the textbook you learned that the density of Fe was 7.8 g/cm^3. What would be the volume of 3.5 g of iron? Show your work.

J. Ways of Separating Matter

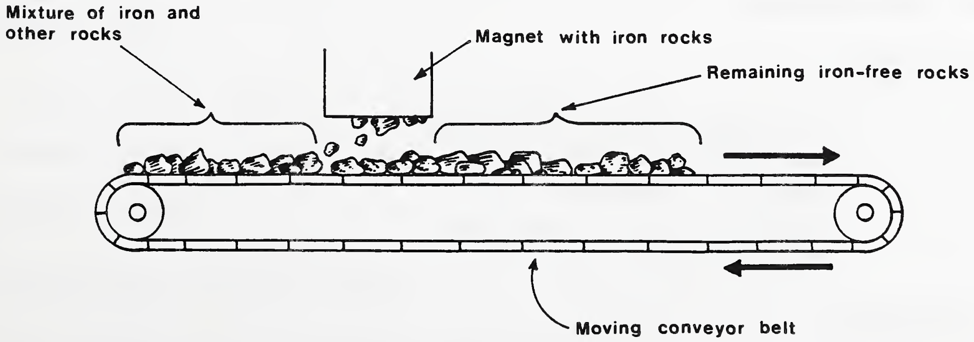
As you learned before, mixtures can be separated by physical means since the various components are not combined chemically. Compounds, however, can be separated into its elements by chemical means.

Various means can be used to separate mixtures depending on their different properties. For example, some differences in physical properties are color, odor, taste, freezing point, magnetic properties, solubility in water, density, hardness, malleability, ductility, luster, electrical conductivity and others. (Always have a dictionary handy for words you don't know the meanings of.)

You can use your own imagination as to how you can separate substances using these different physical properties. Ways in which malleability and magnetism can be used will be discussed below and you will find other ways in the next lab.

Silver is very malleable which means that it can be pounded into thin sheets. If we were to give you twenty small pieces of "rock", about the size of marbles, and told you that eight are silver and the other twelve are just ordinary rock,—assuming that color, density, the outer texture of the rocks and all other physical properties are the same, except for the malleability, how would you separate the mixture? The simplest way would be to hit each rock with a hammer. The silver will be deformed like a piece of putty but the other rocks will tend to splinter up and crack.

Iron, nickel and cobalt along with many compounds containing these and other elements can be separated from other rocks by using their ability to be attracted to magnets. Below, is an illustration as to how a copper mill can separate the iron from its crushed rocks.



Of course, compounds can be separated by chemical means. We won't go into this in great detail. Ways of breaking up compounds can be electrical such as in the electrolysis of water which produces hydrogen and oxygen. Your own body uses enzymes and hydrochloric acid to break down the food you eat into substances it can readily use. For example, table sugar, sucrose ($C_{12}H_{22}O_{11}$), is broken down chemically into glucose ($C_6H_{12}O_6$) which your body can use.

Exercise 5

1. Raw milk contains sugar, fat and water. If you were given a glass of milk and any equipment you needed, how would you separate the fat from the sugar and water?

____

____

____

2. You are given two "silver" dollars, one made of pure silver and the other one made of plastic but coated with a thin layer of silver colored paint so it looked exactly like the silver dollar in every way. Give two physical tests you can use to determine which is which, without damaging any piece in the smallest way. Assume you have all necessary equipment.

(a) ______________________________

(b) ______________________________

K. Experiment 8 SEPARATION OF SAND AND SALT

In the previous section, various ways were discussed on how to separate various mixtures. In this lab, you will actually do a separation of sand and salt.

Materials:

two large beakers or bowls
* mixture of sand and salt
spoon for stirring
balance
thermometer

Procedure:

Weigh the sand and the salt. Total weight is __________. Put the contents into a large container and add water. Stir the solution until you feel no more materials will dissolve in the water. By this time, the water should taste extremely salty. Then pour the liquid into another container and let the remaining sand dry.

Boil the clear salt solution off until the salt is almost dry and let the remaining moist salt dry overnight. Take the temperature of a cup of boiling water as well as the temperature of the salt solution as soon as it begins to boil, when the water is about half gone, and when the water is almost gone. Always be sure the bottom of the thermometer is immersed completely in the water, but does not touch the bottom of container.

For students who do not have our Lab Kit:

Follow the instructions for kit students, but do not taste any chemical or solution for saltiness. Ask your teacher for a mixture of about 40 g of table salt (NaCl) and any other white salt whose solubility product is less than 10^{-6} eg. $CaCO_3$ or $Al(OH)_3$. The lower the solubility product is, the better your results will be. Have someone other than yourself weigh each salt, and record the mass, and then the person should mix the salts together. For best results, mix table salt with ash tray sand if possible.

When dissolving the two salts, an excessive amount of water will dissolve too much of the salt with a low solubility product. Perhaps half a glass of water is all you need in order to completely dissolve just the table salt, and practically none of the other salt. After you have completed the experiment, compare your masses of the two salts with the other person's recorded masses of the two salts.

Observations:

Weight of total mixture ______________

Weight of sand (or other salt) ______________

Weight of salt ______________

Temperature of boiling water ______________

Temperature of initial boiling salt solution ______________

Temperature of intermediate boiling salt solution ______________

Temperature of boiling salt solution near the end ______________

Exercise 6

1. What percentage of your initial mixture was sand and what was salt? (Appendix 11 may help here)

For non-kit students only:

(a) How did the weights of each of the two salts compare with the weights you were given initially?

____

____

____

(b) Can you explain the differences, if any?

____

____

____

2. If the weight of the sand plus the weight of the salt was not equal to the initial weight, how would you explain the difference?

____

____

____

3. Hopefully, every temperature you took was different. How were the solutions with the higher temperature different from the solutions with the lower temperatures?

____

____

____

4. How does the addition of salt affect the freezing temperature of water?

____

____

____

5. Assume the density of salt is 2.17 and the density of sand is 2.60. In what way could you separate the two items if you had the appropriate liquid? (Assuming this liquid does not dissolve salt or sand, what physical property would this liquid have to have to separate the two substances?)

____

____

____

END OF LESSON 4

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ATOMIC THEORY

A. Experiment 9: Black Box Experiment
B. Dalton's Model
C. Gay Lussac and Avogadro
D. Rutherford's Model
E. Periodic Table

A. Experiment 9: BLACK BOX EXPERIMENT

To do: Read Exp. 5-1 in Lab Manual, and see the note in the kit.

For students who do not have our Lab Kit:

If you are a school student, ask your teacher for any black box he may have lying around since this experiment is very common and may possibly have been used in either grades 10, 11 or 12. After you have done the experiment, do not open the box but ask your teacher what the contents were.

If the above is not possible, have someone else insert two or three small objects into a shoe box or other such size box and tape it up after placing in the objects. You are not to see the objects until every effort has been made to try and guess the contents.

Procedure:

Any manipulations which do not involve opening the box and do not damage the box or its contents are permissible.

Fill in the table below, using the hypothesis you are given and add a few of your own. Since a hypothesis is something you can test, test your hypothesis and give an "observation" of what you see, hear, feel or smell. As well, give an interpretation based on your observation. One is given to get you started.

Hypothesis:	Observation: (for each object)	Interpretation: (for each object)
1. The box contains 2 objects	3 distinct sounds are heard when the objects fall to an end.	The box has at least 3 objects.
2. The objects are completely round.	(a) (b) (c)	(a) (b) (c)

Hypothesis:	Observation: (for each object)	Interpretation: (for each object)
3. All objects are the same.		
4. All objects are magnetic.		
5. The objects are very large.		
6. The objects are very heavy.		
7.		
8.		

Exercise 1

1. Based on your interpretation as to what the various objects are like, what is your conclusion about the three objects. If you are not sure what the objects are, at least draw a diagram as to what you think each object approximately looks like.

2. Why is this called a black box experiment?

3. Would you call cures for cancer or the internal structure of the atom as present day black boxes? Explain.

It must be frustrating to try to guess the contents of this box when it would be so simple to open it. However, the whole idea of the experiment was to indicate that scientists often have to go through similar experiences because one can't look into something that easily and come up with a solution, for example a cure for cancer. Hopefully, you realized that much information could be gathered by indirect methods and with a little imagination, you can draw pictures of what you think was in the box. You have probably seen the contents of the box at one time or another. A scientist, however, who is trying to describe an atom, may have a real problem as he is trying to describe something which he has never seen, but from which he has much indirect evidence. One real problem which has not been solved yet is to make a model which adequately describes the complete nature of light. Some experiments indicate that light is made of waves, while other experiments indicate that light is made of particles. At this stage we do not seem to be intelligent enough to come up with a model to explain both observed properties at the same time.

Now open the box and check your conclusions. You can see the contents. Can you think of any other tests that you could have used to give you a more accurate model? If available, what other equipment can one use to identify the objects.

____

____

B. Dalton's Model

To do: Read sections 5-1, 5-2, and 5-3.

Exercise 2

1. List four more major points of Dalton's atomic theory.

   (a) All matter consists of indivisible particles called atoms.

   (b) ____________________________________

   (c) ____________________________________

   (d) ____________________________________

   (e) ____________________________________

2. (a) On the basis of the knowledge we have now, we know that all of Dalton's conclusions were not correct. Do you suppose that he arrived at good conclusions on the basis of the knowledge he did have? ________________ If you don't think so, give reasons.

   ____

   ____

   ____

   (b) Do you suppose we have any theories today which explain everything we know about the atom at the present time, but may need revision if completely unexpected experimental results show up? __________

3. With which of the following statements would you agree?

   (a) Dalton should not have proposed that oxygen is eight times heavier than hydrogen unless he was absolutely certain of the fact. Otherwise, he might lead other scientists down the wrong track.

(b) Proposing this wrong prediction was a good thing since other scientists could use this prediction which seemed reasonable enough. They could try to come up with new experiments which would either prove or disprove Dalton's hypothesis.

_______________ (a or b?)

4. How would Dalton have explained the existence of the compounds NO and NO_2 ?

____

____

____

____

5. List four ways in which Dalton's work helped future scientists.

(a) ____

____

(b) ____

____

(c) ____

____

(d) ____

____

C. Gay Lussac and Avogadro

To do: Read section 5-4.

Gay Lussac formulated the Law of Combining Volumes which states that at a given temperature and pressure, gases combine in volumes bearing a simple ratio to one another and to the volumes of the gaseous products.

Avogadro states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules. It turns out that 22.4 L of any gas at 1 atmosphere (100 kilopascal)(kPa) and at 0°C contains 6.02×10^{23} molecules. Since the atomic weight of argon is 40.0, 6.02×10^{23} molecules of argon weigh 40.0 g.

D. Rutherford

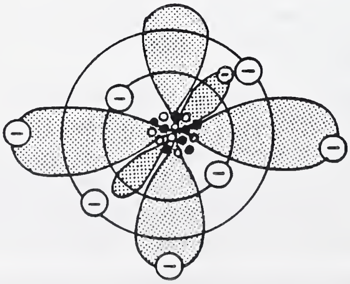
To do: Read sections 5-5 to 5-8

The structure of an atom consists of a dense, positively charged nucleus surrounded by electrons. Electrons are light, negatively charged particles, whereas a proton is a dense, positively charged particle. Every neutral atom has as many electrons as protons. The number of electrons or protons which any element has is given by its atomic number. For example, the atomic number of oxygen is 8, so each oxygen atom has 8 electrons and 8 protons. The atomic weight is equal to the total number of protons plus neutrons. A proton weighs about as much as a neutron and both are much heavier than the electron. Since the atomic weight of oxygen is 16 and it has 8 protons, how many neutrons does oxygen have? Hopefully you guessed 8, because the

$$\text{atomic weight} - \text{atomic number} = \text{number of neutrons}$$

Both the neutrons and the protons are contained in the nucleus and occupy very little volume. The lightest twenty elements often have as many protons as neutrons, but after this, you have more neutrons than protons. Since positive charges repel, can you imagine what would happen if you had twenty protons together without any neutrons to kind of keep the protons together? Very strong nuclear forces keep the nucleus together.

When looking at the table of atomic weights inside the front cover of the textbook, notice that all atomic numbers are whole numbers but the atomic weights are often fractions. The reason for this is that all elements of a given atom have a specific number of protons but the number of neutrons may vary slightly. For example, a chlorine atom ALWAYS has 17 protons in the nucleus otherwise it wouldn't be chlorine. However, the chlorine nucleus may have 18 or 20 neutrons. The atomic weight averages out to 35.5 because there are more atoms with 18 neutrons than with 20 neutrons. No atom has a fraction of a neutron.



The above diagram shows an extremely simplified sketch of an oxygen atom with 8 central protons, 8 central neutrons, and 8 orbiting electrons. The above diagram is essentially Rutherford's model of the atom, based on his experiments.

From the above sketch, the protons and neutrons are seen to be packed into a dense nucleus, and for our purposes the protons can be considered immobile. Electrons, however, can easily be transferred from one atom to the next. An atom can become negatively charged by gaining electrons, and it can become positively charged by losing electrons which would result in the atom having more protons than electrons. Any neutral atom that has gained or lost electrons is an ion. In other words, an ion is a particle that is either positively or negatively charged. The neutral chlorine atom has 17 protons and 17 electrons. The choloride ion (Cl^-) has 17 protons but 18 electrons.

Because electrons are so mobile, it is the electrons which are responsible for all chemical behaviour. Also electrons are responsible for the conduction of electricity in a wire.

To do: Practice Exercise problems 3, 4, 5, and Self Test problems 3, 4, 5, 10 (page 151)

After you have corrected your answers on the exercises above, do the exercise below.

Exercise 3

1. Match the name with the appropriate line at the right.

A. Becquerel
B. Thomson
C. Rutherford
D. Gay-Lussac
E. Roentgen
F. Chadwick
G. Dalton
H. Soddy and Richards
I. Millikan
J. Avogadro
K. Democritus

__________ discovered the neutron
__________ stated that the atom has a dense, positively charged nucleus
__________ discovered isotopes
__________ discovered radioactivity
__________ discovered X rays
__________ discovered electrons
__________ found charge on electrons
__________ suggested Law of Combining Volumes
__________ equal volumes of gases have equal numbers of molecules
__________ provided a basis for predicting possible formulas for compounds
__________ introduced the atom

2. Of the five major points in Dalton's atomic theory, which do you now know to be false and give a reason why.

3. Briefly, give three reasons for studying the development of atomic theory in historical context.

(a)

(b)

(c)

4. Briefly describe Rutherford's experiment after which he decided that the atom consists of a dense positively charged nucleus surrounded by electrons.

E. Periodic Table

The periodic table combines chemical and physical properties, and atomic structures.

Study the periodic table on pages 170 and 171. Notice that the atomic numbers increase from left to right. Also note that about four-fifths of all elements are metals, and that one-fifth, the upper right hand corner, are non-metals. The alkali metals, the alkaline earth metals, and the noble gases are all in a vertical column or group. The vertical columns are identified by Roman numeral group numbers which range from IA to VIII A across the top and from IB to VIII B across the middle. The B groups are the transition metals and will not concern us much now. All elements in each A group have similar physical and chemical properties. (Even though the properties are similar in a given group, the elements always become more metallic as you go down a particular group. In group IV A, you have the extreme case where you go from a non-metal (C) to a metal (Pb).)

As you learned in the previous section, the atomic number refers to the number of protons or electrons in the neutral atom. The number of protons has very little to do with chemical behaviour, but the structure and number of the electrons is extremely important in determining chemical behaviour. Out of the whole mass and volume of an atom, it is only the outer electrons which are completely responsible for all chemical behaviour. What is unique about each group or vertical column is that the outer electrons are arranged in a similar manner, resulting in similar physical and chemical properties.

Exercise 4: Noble Gases

Much of what is covered in Chapter 6 of the text is beyond the scope of this course. What we would like you to do now is read section 6-6. As you read it, pay particular attention to the details that follow below. Answer questions where required by filling in the blanks or by underlining the correct word or words.

1. Note that the noble gases, with the exception of helium, have eight outer electrons (page 165). Their group number is VIII. Is this a coincidence?

   ____

2. Don't worry about what ionization energy is. Note that the ionization energy varies from 24.5 to 10.7 and it goes down as you go down the group of noble gases.
3. Note the melting points and how they vary with noble gases.
4. The density for noble gases is given in grams per litre since they are all gases. Densities (increase or decrease) down the group.
5. Atomic radii (increase or decrease) down the group. (Underline correct word)

Exercise 5: Alkali Metals

1. How many outer electrons does each alkali metal have? __________

   What is their group number? __________

2. Ionization energies (increase or decrease) as you go down the group, but all ionization energies are (higher or lower) than for noble gas.

3. Melting points (increase or decrease) as you go down the group and they are (much higher or much lower) than the melting point for noble gases.

4. The density for alkalis is given in (   ), which means the alkali metals are roughly 1000 times denser than the noble gases. Densities (increase or decrease) down the group.

5. Atomic radii (increase or decrease) down the group.

Exercise 6: Halogens

1. How many outer electrons does each halogen have? __________

   What is their group number? __________

2. Ionization energies (increase or decrease) as you go down the group.

3. The ionization energies for halogens are very close to (ionization energies for noble gases or ionization energies for alkalis).

4. Note that there is a great difference in melting points and densities of the halogens since the top ones are gases, bromine is a liquid, and iodine is a solid at room temperature.

5. Densities (increase or decrease) as you go down the group.

6. Atomic radii (increase or decrease) down the group.

Hopefully, the above questions and comments made it clear that various groups have certain properties and that the properties usually follow definite patterns as you go down the group. Other chemical and physical properties follow similar patterns for the various groups in the periodic table.

The properties of the alkaline earth metals are not given, but try to predict the properties of the alkaline earth metals on the basis of what you have learned.

Exercise 7: Alkaline Earth Metals

1. What is the group number of the alkaline earth metals? __________
Therefore they have __________ outer electrons.

2. Ionization energies (increase or decrease) as you go down the group.

3. The ionization energies of the alkaline earth metals are close to those of (the alkali metals or the halogens).

4. Melting points (increase or decrease) as you go down the group.

5. Melting points would be close to that of (alkalis or noble gases).

6. Densities (increase or decrease) as you go down the group.

To do: Read section 6-7

Exercise 8

1. Who initially arranged elements according to atomic weight?

____

2. Who then arranged them by atomic number? __________________________

3. In most cases, it doesn't make much difference how the elements are arranged, since the atomic weight and atomic number usually go up at the same time. However, there are a few exceptions. One of these exceptions is argon and potassium. Which of these two has the higher atomic weight? ______________ Which has the higher atomic number?

____

4. State the periodic law. ____

____

____

____

To do: Read section 6-8

5. Give three trends which can be noted regarding elements of the periodic table.

(a) ____________________

(b) ____________________

(c) ____________________

END OF LESSON 5

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(2) All exercises are completed. If not, explain why.
(3) Your work has been re-read to ensure accuracy in spelling and lesson details.
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Send all letters in a separate envelope.

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Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope. Correspondence lessons will travel faster if first-class postage is used.**

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

CHEMICAL EQUATIONS

A. IUPAC Rules
B. Mole Concept

A. IUPAC Rules

Writing and Naming Chemical Compounds

In Lesson 4, chemical symbols were discussed. A symbol of one or two letters can be used to indicate any element. Unfortunately, life is not so simple that everything is just composed of pure elements. Most things you will meet are compounds which consist of one or more elements joined together. To represent a compound, you will have to write the appropriate symbols together. For example, one atom of carbon (C) combines with one atom of oxygen (O) to form carbon monoxide (CO.) It is also possible that one atom of carbon can combine with two atoms of oxygen to form carbon dioxide (CO_2). Note that when you have two atoms of any element, the 2 is written as a subscript, slightly below and to the right of the element indicated.

What is mentioned in the paragraph below is in reference to section 8-3, which you will be asked to read.

Due to the way this course is set up, chemical bonding, which is covered in Chapter 7 of the textbook, is not covered in Chemistry 10. When any reference is made to any types of bonding, or any other material taken from earlier chapters, just ignore these sections. However, the bottom portion of page 206 is very important. Also, pay particular attention to the portion on diatomic molecules and monotomic molecules. There are only 8 diatomic molecules. They are: hydrogen (H_2), nitrogen (N_2), oxygen (O_2), and the five halogens of group VIIA: F_2, Cl_2, Br_2, I_2, At_2. Whenever one of these eight elements combines with anything else, always write these elements in diatomic form when they appear alone, never as a single atom. However, when these elements appear in a compound they are never in a compound as a diatomic gas. The following illustration should make this clear. When magnesium (Mg) reacts with oxygen, (O_2) magnesium oxide (MgO) is produced. To write this with symbols, it would be

$$Mg + O_2 \rightarrow MgO$$

Do <u>not</u> write:

$Mg \quad O \rightarrow MgO, \times$
nor $Mg \quad O_2 \rightarrow MgO_2 \times$

To do: Read section 8-3 and 8-4

The oxidation numbers on the periodic table may look very confusing at first, but the situation is not nearly that bad. For one thing, most of the elements will never be used at this stage and there are some general guidelines for the other elements which you will use. The last one-third page of 206 pretty well summarizes most of what you need to know about oxidation numbers of various elements. Look at the elements in the 'A' groups of the periodic table. The elements in group VIII A are happy with the 8 electrons they have in their outer ring. Inner electrons have no effect on an elements oxidation number. Therefore, when an element has eight outer electrons, the structure of all of its electrons is the same as that of the closest inert gas. For example, potassium (K), whose atomic weight is 19, has 19 electrons in the neutral atom (and also 19 protons). Argon, the closest inert gas, has 18 electrons and 18 protons. To have an electronic structure like argon, potassium must lose 1 electron, leaving it with 18 electrons, but 19 protons. Since it then has one more proton than electron, the charge of the potassium ion is +1. Using similar reasoning, all elements in group I A when they become ions, have a charge of +1. Elements in group IIA have two more electrons or negative charges than the nearest inert gas, so when an element from group IIA becomes an ion its charge is +2.

An element like chlorine, however, would have to lose 7 electrons if it were to be like neon. It is much easier for chlorine to gain one electron and become like argon, rather than lose seven electrons. In gaining one electron the charge is -1. On the right side of the periodic table, the oxidation number would be 8 minus the group number. For example, nitrogen is in group VA; its oxidation number is 8 - 5 = 3. However, the elements on the right gain electrons, which means their oxidation number is negative. So the oxidation number for nitrogen is -3.

Below is a brief summary of what is covered above.

Group	I A	II A	III A	IV A	V A	VI A	VII A	VIII A
common oxidation number	+1	+2	+3	+4	-3	-2	-1	

Keep in mind that the above table is a greatly simplified version of the table on page 208, however, it should be sufficient for our purposes.

In addition to knowing the oxidation numbers of the 'A' group elements, you will also have to know the oxidation numbers of a few transition metals. Transition metals, for example chromium, copper, iron are a bit tricky in that they often have two common oxidation numbers. Also, some of the 'A' group metals, for example tin and lead also have two common oxidation states.

In addition to elements, certain groups of atoms act like a single element in a chemical reaction and they have their own particular oxidation numbers which denotes the charge for that group. For a list of common elements and ions you will encounter, please see Appendix 6, page 295 in the textbook.

Of course, the whole purpose of knowing all of these oxidation numbers is that you will be able to write the formulas for various compounds. The sum of the oxidation numbers in any compound must always be zero. For example, if aluminum Al^{+3} reacts with chlorine Cl^{-1}, you need three chlorine ions for each aluminum ion since $+3 + 3(-1) = 3 - 3 = 0$. Another way of saying the same thing is that aluminum must give away 3 electrons to become like the nearest inert gas, but each chlorine atom can only accept 1 electron to become like its nearest inert gas. Obviously, you need 3 chlorine atoms to accept all three electrons from the aluminum. Thus the formula for aluminum chloride has to be $AlCl_3$.

Exercise 1

1. The following elements are in the atmosphere: oxygen, argon, krypton and nitrogen. Give the symbols for 1 mole of each element. (Hint: some elements are diatomic; others are monotomic.)

________ ________ ________ ________

2. Using a periodic table (page 170 & 171 or back cover) find the group number of the following elements, and then, using this information, indicate what charge the ion of that element would have.

Na ________, B B^{+3}, Al ________, Mg ________,

Ca ________, F ________, Cl ________, O O^{-2}

S ________,

3. Using appropriate symbols, write the reaction when calcium combines with oxygen to produce calcium oxide.

____

Exercise 2

Following the three steps on page 207, please write the formula for the compounds when the following elements are combined (see the next page.) Omit step 4 for now. Two have been done for you. Note that when neither oxidation number is one, both subscripts might have to be changed.

	Step 1	Step 2	Step 3 (formula)	Step 4 (name)
(a) chlorine + sodium	Cl^{-1}, Na^{+1}	Na Cl	NaCl	sodium chloride
(b) chlorine + magnesium				
(c) fluorine + aluminum				
(d) oxygen + sodium				
(e) oxygen + magnesium				
(f) nitrogen + lithium				
(g) oxygen + aluminum	O^{-2}, Al^{+3}	Al O	Al_2O_3	aluminum oxide
(h) sulfur + aluminum				
(i) magnesium + nitrogen				
(j) chlorine + copper (Cu^{+1})				
(k) chlorine + copper (Cu^{+2})				
(l) fluorine + tin (Sn^{+2})				
(m) fluorine + tin (Sn^{+4})				
(n) sodium + acetate* ($C_2H_3O_2^{-1}$)				omit

* Treat the acetate ion like a single element

To do: Read section 8-5.

Most of this section is very straightforward. To name binary compounds (compounds with only two different elements) use the italicized names on page 211.

Some difficulty may arise with compounds which contain two non-metals. This is partly due to earlier things not covered in the textbook and due to the oversimplication that was made on page 3 of this lesson. However, it is easy to get around this difficulty without too many problems. You will notice that in a binary compound, the metal is always named first and the non-metal is always named last. Metals are on the left of the periodic table and non-metals are on the right. So when you have two non-metals forming a compound, name the element which is farthest to the left in the periodic table first, and the element which is closest to the right in the periodic table should be named last. To be more specific, whether you are dealing with a metal and a non-metal or with two non-metals, always check which element is closest to fluorine. This element is then named lastly in a binary compound.

Exercise 3

1. Name all of the compounds from the previous exercise (#2). This should be very straightforward since the first element that you should have written in step 3 is the first one you name. Don't forget the "---ide" ending for non-metals. (Use preferred names on page 211.)

2. Name the following compounds.

(a) PCl_3

(b) PCl_5

(c) NO_2

(d) SO_2

(e) CO

(f) CO_2

(g) P_2O_3

(h) P_2O_5

To do: Read section 8-6.

As before, the component with the positive oxidation number is named first. An important difference, however, is that the ending of complex ions is not changed.

When more than one complex ion is present in a molecule, always put brackets around the ion and then use a subscript to indicate how many of a particular ion you need. For example, sodium hydroxide is NaOH, but calcium hydroxide must be written as $Ca(OH)_2$. $Ca(OH)_2$ implies one calcium atom and two groups of OH. It is wrong to write $CaOH_2$ since this implies the presence of one calcium atom, one oxygen atom, and two hydrogen atoms. Brackets are essential here. It is also wrong to write CaO_2H_2 because of the way the molecule is arranged, — you have two distinct OH groups. As an analogy, two boys who weigh 45 kg each are not the same as a single boy who weighs 90 kg. It is also wrong to write Ca2OH. Subscripts must always be used to indicate two or more complex ions in a given molecule.

Along a similar line, when you talk about a diatomic molecule of hydrogen, you must write H_2. If you write 2H, you are implying 2 individual atoms of hydrogen.

To do: Problems #1 and #2 on page 221, and #2 and #3 on page 222. Check your work.

When you feel you understand this material, please do the exercise below.

Exercise 4

1. Write the names for the following compounds. (Refer to Table of Complex Ions)

(a) NH_4OH ____________________

(b) $Ca(OH)_2$ ____________________

(c) Na_2SO_3 ____________________

(d) Na_2SO_4 ____________________

(e) $Al_2(CO_3)_3$ ____________________

(f) $AlPO_4$ ____________________

(g) $CaCl_2$ ____________________

2. Write the formulas for the following compounds.

(a) ammonium carbonate ____________________

(b) calcium acetate ____________________

(c) rubidium nitrate ____________________

(d) magnesium sulfate ____________________

(e) magnesium sulfite ____________________

(f) magnesium sulfide ____________________

(g) aluminum carbonate ____________________

(Appendix 4 summarizes the naming of compounds and Appendix 5 covers formula writing.)

B. The Mole Concept

To do: Read sections 8-1 and 8-7 in the textbook.
(In SI, the abbreviation for mole is mol.)

In summary, it is impossible to weigh individual atoms. Consequently, a large number of atoms are weighed and from this information the exact number of atoms can be calculated. Inside the front cover, you have atomic weights for all elements. The atomic weight of oxygen is 16.0. On an atomic scale, this means that each oxygen atom has a total of 16 particles in the nucleus. It turns out that oxygen has 8 neutrons and 8 protons. On a laboratory scale, however, it means that 6.02×10^{23} atoms of oxygen weigh 16.0 g. One mole of anything has 6.02×10^{23} particles, so a mole of oxygen atoms weigh 16.0 g.

In section 8-7, a differentiation is made between formula weights and molecular weights. To simplify things at this stage, just molar masses will be covered. For example, the molar mass of carbon tetrachloride is 154 g/mol and the molar mass of potassium iodide is 166.0 g/mol.

From the preceding discussions in the textbook, you should be able to give the weight of a mole of atoms of any substance. (Just read off the atomic weight in grams.) For example, the weight of a mole of copper atoms is 63.5 g. You should also be able to give the molar mass of every compound. (Just add the weights of the individual atoms together.) For example, the molar mass of carbon dioxide (CO_2) is $12.0 + 16.0 + 16.0 = 44.0$ g.

When you are asked for the molar mass of any diatomic substance, you must double the atomic weight. For example, the atomic weight of chlorine (the weight of a mole of atoms) is 35.5 g. But the molar mass of chlorine (the weight of a mole of molecules) is 35.5 + 35.5 = 71.0 g, since a molecule of chlorine is diatomic, its formula is Cl_2.

If you are asked for the molar mass (ie. the weight of one mole) of any substance, just add the atomic weights of the individual atoms together. However, in many cases you may be interested in the weight of 0.5 mol of a given substance, or you may be interested in the weight of 3 mol of a given substance. What you do then, is not to panic, but simply add the weights of the individual atoms together, and multiply by the appropriate number of moles. For example, the molar mass of one mole of potassium iodide is 166.0 g. A half of a mole would weigh 166.0/2 = 83.0 g. Three moles would weigh 166.0 × 3 = 498.0 g.

You may also be asked to find how many moles of KI there are in 8.3 g of the substance. Obviously, this is much less than one mole, so be sure your answer is reasonable and that you have the right number on top and the right number at the bottom in your division. Use this formula:

$$\text{number of moles} = \frac{\text{weight of substance}}{\text{weight of one mole of substance}}$$

We find that in above case,

$$\text{number of moles} = \frac{8.3\ \text{g}}{166.0\ \text{g/mol}} = 0.050\ \text{mol},$$

or 5.0×10^{-2} mol.
(Check: is this less than 1?)

By the way, how many molecules would be in 8.3 g of KI?

1 mol has 6.02×10^{23} molecules
0.05 mol has X molecules

By cross multiplying, $1X = 6.02 \times 10^{23}$ times 0.05
$X = 3.01 \times 10^{22}$ molecules

To do: Do problem #3 in the Practice Exercise on page 221, and #1 in the Self Test on page 222. If you feel you understand this after checking your answers, do the exercise on the next page.

Exercise 5

1. Calculate the molar masses of the following substances. Be accurate to two decimal places.

   (a) $Ca(NO_3)_2$

   (b) $HgCl_2$

   (c) $(NH_4)_2SO_4$

   (d) Mn_2O_5

   (e) $Fe(C_2H_3O_2)_3$

   (f) CaO

   (g) $KMnO_4$

   (h) O_2

   (i) O_3 (O_3 is ozone, which contains 3 atoms of oxygen per molecule)

2. The weight of a proton or neutron is 1.66×10^{-24} g. Since the atomic weight of hydrogen is 1.0, it means that each hydrogen atom contains only one proton but no neutrons. (The weight of the electron is negligible.) How many hydrogen atoms would be needed to make 1.0 g of hydrogen?

$$\text{HINT:} \quad \frac{\text{X atoms}}{\text{1 atom}} = \frac{1.0\ \text{g}}{1.66 \times 10^{-24}\ \text{g}} \qquad \text{Find X}$$

3. One mole of any substance has 6.02×10^{23} molecules.
   (a) How many molecules are there in 2 mol of any substance?

   (b) How many molecules are there in 0.5 mol of any substance?

   (c) One mole of $MgSO_4$ has 6.02×10^{23} molecules. Each molecule of $MgSO_4$ has 6 atoms. Is this right? ________ How many atoms are there in 1 mol of $MgSO_4$?

   (d) How many molecules are there in 0.01 mol of $NaOOCC_{17}H_{35}$? (This is the formula for a soap molecule.)

End of Lesson 6

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CHEMICAL EQUATIONS - (continued)

C. Writing Chemical Equations
  (1) Qualitative
  (2) Quantitative
D. Balancing of Chemical Equations
  (1) Using Moles of Atoms
  (2) Using Atoms
  (3) Conservation of Mass

C. Chemical Equations

Chemical equations may be either (1) qualitative or (2) quantitative depending on the information that is available.

(1) Rules to follow in writing a qualitative chemical equation. It is very important to know what is reacting (reactants) and the substances produced in the reaction (products). Also one must know the correct formulas for the reactant(s) and for the product(s) before writing the equation. As a rule, the formula for the reactants appears on the left side and the formula for the products on the right side of the arrow.

At this point it is not necessary to know how many moles or how many grams of reactants take part in the reaction, nor is it important to know how much of the product is formed in the reaction. These quantities will be dealt with later in the lesson. Also we will not worry about balancing equations now, therefore coefficients will be ignored in the reactions that follow. However, gases such as O_2, H_2, and N_2, as well as the halogens, are diatomic.

You may be quite aware of the many chemical reactions taking place about you, but you really had no opportunity to express them in equation form. To think of a few examples, consider:

(a) Silver cutlery tarnishes in your home. Why does this occur?

Silver combines with oxygen in the atmosphere to produce a dark coating of silver oxide. The equation for this reaction is:

$$Ag + O_2(g) \rightarrow Ag_2O$$

(b) Charcoal (carbon) burns in your barbeque to produce carbon dioxide. Equation:

$$C + O_2(g) \rightarrow CO_2(g)$$

(c) A candle burns on your birthday cake to produce carbon dioxide and water. Equation:

$$C_{25}H_{52} \text{ (paraffin)} + O_2(g) \rightarrow CO_2(g) + H_2O$$

(d) Baking soda neutralizes an acid stomach.
Equation:

$$NaHCO_3 + HCl \text{ (acid in stomach) } \rightarrow H_2O + CO_2(g) + NaCl$$

The above reactions are examples of qualitative chemical equations.

Exercise 1

Using the rules given in this lesson for writing equations and using examples above, please write equations for the following reactions: Check list of diatomic molecules on page 1, Lesson 6.

1. A copper cooking utensil tarnishes in air to produce a dark coating of copper II oxide. (CuO)

____

2. Gasoline (C_8H_{18}) burns completely in your automobile engine to produce carbon dioxide and water.

____

3. Iron rusts to produce iron II oxide.

____

4. Sodium metal reacts with water to produce hydrogen and sodium hydroxide.

____

5. Potassium combines with bromine to yield potassium bromide.

____

6. In flash bulbs, magnesium burns in oxygen to produce a white solid magnesium oxide.

____

7. Nitrogen combines with hydrogen (under special conditions) to produce ammonia (NH_3).

8. Chlorine reacts with sodium to produce common table salt.

Qualitative chemical equations may also show energy as part of the reaction. For example, water may be decomposed into hydrogen and oxygen using electrical energy.

$$H_2O(\ell) + \text{energy} \rightarrow H_2(g) + O_2(g)$$

The above represents an <u>endothermic</u> reaction because energy is <u>absorbed</u> during reaction.

A reaction may show energy is <u>released</u> as in burning charcoal. We obtain energy in terms of heat.

$$C + O_2 \rightarrow CO_2 + \text{energy}$$

The above reaction illustrates an <u>exothermic</u> reaction. Again we shall not discuss the <u>quantity</u> of energy absorbed or released since we are dealing with only qualitative aspects of a chemical reaction.

<u>Exercise 2</u>

Write an equation to show decomposition of potassium chlorate ($KClO_3$) by heating it strongly to produce oxygen and potassium chloride.

(2) <u>Rules to follow in writing Quantitative chemical equations.</u> Here again you will be expected to know the correct molecular formula for the reactants and products but in addition you will have to show that (1) atoms are conserved and that (2) mass is conserved.

To meet these requirements you will have to <u>balance</u> your equation. Balancing of equations involves a simple bookkeeping device in which one keeps track of atoms taking part in a chemical reaction. The number of atoms entering a reaction must equal the number of atoms forming the new products. One may count these atoms by using the convenient mole concept. In other words, we will count moles of atoms in a given chemical reaction.

If you are given the number of moles of reactants and the number of moles of products then it would be quite simple to count the number of moles of atoms reacting and the number of moles of atoms (of each element) produced by the given reaction. For example: you may be given this:

> Two moles of red mercuric oxide are heated to produce one mole of oxygen and two moles of mercury. The above can be written as a chemical equation: $2\ HgO + \text{energy} \rightarrow O_2 + 2\ Hg$

Is it possible to determine the number of moles of atoms reacting or the number of moles of atoms produced by observing the above equation?

Details regarding this procedure will be discussed in the next topic under Balancing of Equations.

<u>Exercise 3</u>

Given that one mole of propane (C_3H_8) reacts with five moles of oxygen to yield three moles of carbon dioxide and four moles of water, write the chemical equation for this reaction showing the <u>number of moles</u> of each compound involved.

D. <u>Balancing of Chemical Equations</u>

(1) For the balancing of equations using <u>moles of atoms</u>, study the examples below and refer to Appendix 7 pages 296-297 of your textbook.

Magnesium combines with oxygen in the air to produce a white solid, magnesium oxide. To write the equation, you must know the formula for magnesium, oxygen and magnesium oxide.
Therefore you will have:

$$Mg + O_2 \rightarrow MgO$$

In order to conserve moles of atoms we must find numerical coefficients to place <u>before</u> each formula to obtain the same number of moles of atoms of each element on the left and on the right of the arrow. That is, moles of magnesium atoms on the left side must equal moles of magnesium atoms on the right side of the arrow. The same must be true for all other elements. The process of finding these coefficients is called <u>balancing the chemical equation</u>.

Suppose one mole of oxygen gas is consumed in the reaction.

$$Mg + 1O_2 \rightarrow MgO$$

But one mole of oxygen gas contains two moles of oxygen atoms. On your right side, one mole of magnesium oxide contains only one mole of oxygen atoms. To balance oxygen, place a 2 in front of MgO which will now give two moles of oxygen atoms.

$$Mg + O_2 \rightarrow 2MgO$$

(With no coefficient, a 1 is understood to be in front.)

But now magnesium is not balanced. Place a 2 in front of Mg. Are all atoms conserved now? ________

$$2Mg + O_2 \rightarrow 2MgO$$

You can use fractions to balance equations. The same equation is correctly balanced as shown below:

$$Mg + 0.5O_2 \rightarrow MgO$$

Are moles of atoms balanced? ___________

SUMMARY

$$2\ Mg + O_2 \rightarrow 2\ MgO$$

$$Mg + 0.5\ O_2 \rightarrow MgO$$

Either way of writing the balanced equation for the burning of magnesium is correct.

If you wish to eliminate fractions you may multiply the second equation by 2.

$2 \times (1\ Mg + 0.5\ O_2 \rightarrow 1\ MgO) =$

$2\ Mg + 1O_2 \rightarrow 2\ MgO$ to obtain the first equation.

Please Note: In writing chemical equations we often drop the coefficient 1, but it is never wrong to write it in the equation.

(2) For the balancing of equations using atoms, study the examples below.

Given that hydrogen burns in air to yield water, your equation for the above reaction is as follows:

$$H_2 + O_2 \rightarrow H_2O$$

1 molecule of hydrogen gas or 2 atoms of hydrogen	+	1 molecule of oxygen gas or 2 atoms of oxygen	→	1 molecule of water which contains 2 atoms of hydrogen plus one atom of oxygen

To correct the deficiency of oxygen atoms on the right side of your equation you may place the coefficient 2 before the formula H_2O.

Now you have:

$$H_2 + O_2 \rightarrow \underline{2}\ H_2O$$

In 2 molecules of water there are 2 atoms of oxygen and 4 atoms of hydrogen. To correct the deficiency of hydrogen atoms on the left side you use the coefficient 2 before the formula H_2.

Now you have:

$$\underline{2}\ H_2 + O_2 \rightarrow 2\ H_2O$$

Make a final check to see whether atoms are conserved.

Left side	element	Right side
$2 \times 2 = 4$	H	$2 \times 2 = 4$
$1 \times 2 = 2$	O	$2 \times 1 = 2$
Total $\boxed{6}$		Total $\boxed{6}$

Atoms are conserved!

Exercise 4

Balance the following equation:

$$O_2 + Cl_2 \rightarrow Cl_2O$$

In balancing atoms of oxygen, the equation becomes

$$O_2 + Cl_2 \rightarrow \underline{\hspace{1em}}\ Cl_2O$$

Left side	element	Right side
1 × 2 = 2	O	2 × 1 = 2

The equation requires correction for chlorine atoms.

The balanced equation becomes:

$$O_2 + \_\_Cl_2 \rightarrow \_\_Cl_2O$$

Are all atoms conserved? ________ If they are, show proof below.

Left side	element	Right side
	O	
	Cl	
Total ___		Total ___

Exercise 5

Acetylene burns in air to form carbon dioxide and water, represented by the following unbalanced equation:

$$C_2H_2 + O_2 \rightarrow CO_2 + H_2O$$

First balance carbon atoms by placing the correct coefficient before the formula CO_2.

$$C_2H_2 + O_2 \rightarrow \_\_CO_2 + H_2O$$

Next balance oxygen atoms by placing the correct coefficient before O_2 (fractions may be used).

$$C_2H_2 + \_\_O_2 \rightarrow \_\_CO_2 + H_2O$$

To clear the fraction multiply the equation by ________. The equation becomes: ____.

Are all atoms conserved? ____________________

In summary, please read Section 8-9 (page 216 of the textbook, — rules 1 to 4) on how to write a correct chemical equation.

To do:

Practice Exercise

Textbook - Page 221 - problem 6 (a-e)
Do not send in for correction.
Check your work with answers in your textbook.

Exercise 6

Balance the following chemical equations:

(a) ___C_2H_2 (acetylene) + ___O_2 → ___CO_2 + ___H_2O

(b) ___HgO → ___Hg + ___O_2

(c) ___Cu + ___S_8 → ___CuS

(d) ___Cl_2 + ___NaI → ___I_2 + ___NaCl

(e) ___NaOH + ___H_2SO_4 → ___H_2O + ___Na_2SO_4

(f) ___Li + ___$AlCl_3$ → ___Al + ___LiCl

(g) ___$FeCl_3$ + ___KOH → ___$Fe(OH)_3$ + ___KCl

(h) ___C_3H_8 + ___O_2 → ___H_2O + ___CO_2

(i) ___NaCl → ___Na + ___Cl_2

(j) ___K + ___H_2O → ___H_2 + ___KOH

(3) Conservation of Mass

A balanced equation also shows that mass is conserved. During a chemical reaction mass is neither created nor destroyed but is changed from one form to another. To illustrate that mass is conserved in a chemical reaction, consider the balanced equation for the buring of magnesium.

2 Mg	+	O_2	→	2 MgO
2 mol of magnesium, Mg	+	1 mol of oxygen, O_2	→	2 mol of magnesium oxide, MgO
2 mol of Mg = (2 mol) (24.31 g/mol) = 48.62 g		1 mol of O_2 = 2 mol of O = (2 mol) (16.00 g/mol) = 32.00 g		2 mol of MgO = (2 mol) (24.31 g/mol + 16.00 g/mol) = 80.62 g

48.62 g + 32.00 g = 80.62 g = 80.62 g

The mass of the reactants equals the mass of the products. This shows that not only atoms are conserved but also mass is conserved.

Using your balanced equations in Exercise 6 you can show that mass is conserved. Work through the example below before going on to the next exercise. Consider equation (e) from Exercise 6.

$$NaOH + H_2SO_4 \rightarrow H_2O + Na_2SO_4$$

If the equation is balanced properly, the mass of the reactants should equal the mass of the products.

Reactants		Products	
1 mol of NaOH	= 40.00 g	1 mol of H_2O	= 18.02 g
1 mol of H_2SO_4	= 98.08 g	1 mol of Na_2SO_4	= 142.04 g
total	138.08 g	total	160.06 g

It is quite obvious that mass is not conserved so the equation is not balanced. Balance the equation:

_____ NaOH + _____ H_2SO_4 → _____ H_2O + _____ Na_2SO_4

Now complete the outline below:

Reactants Products

_____ mol NaOH = _____ g _____ mol H_2O = _____ g

_____ mol H_2SO_4 = _____ g _____ mol Na_2SO_4 = _____ g

Total mass of Reactants = _____ Total mass of products = _____

Are you convinced that mass is conserved? _______

Exercise 7

Verify that equations h, i and j in Exercise 6 are correctly balanced by showing their mass is conserved.

Write the balanced equation first, then show your work below.

1. Equation (h):

2. Equation (i):

3. Equation (j):

Exercise 8

In a balanced reaction:

1. Is it necessarily true that moles of atoms are always conserved? Explain.

2. Given $2HgO_{(s)} \rightarrow 2Hg_{(l)} + O_{2(g)}$

Is it necessarily true that moles of molecules are always conserved? Explain.

END OF LESSON 7

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CHEMICAL EQUATIONS - (continued)

E. Using Equations in Problem Solving
F. Types of Chemical Reactions
  (1) Synthesis Reaction
  (2) Decomposition Reaction
  (3) Rearrangement or Replacement Reaction
  (4) Metal-ion Reaction
  (5) Ion-ion Reaction
G. Experiment 10: A Quantitative Study of a Chemical Reaction

E. Using Equations in Problem Solving

Now that you have learned how to write and balance chemical equations you will be able to use chemical equations in problem solving. Since chemistry is a quantitative science, a chemist wishes to know more than the qualitative fact that a reaction occurs; he must also answer questions beginning: "How much ---" or "How many ---". The quantities are usually expressed in grams or moles as will be evident in the sample problems. In order to be successful in this type of work, a thorough understanding of the mole concept is essential. (Please review this concept if necessary). This quantitative relationship implied in a chemical reaction is called stoichiometry.

To do: Read Section 8-8 pages 214-215 and Rules 5 to 8, page 216 of your textbook.

Study examples on pages 216-217 of your textbook.

For a better understanding of this material, please study the examples that have been worked out for you below.

Example 1

Suppose we wish to know how many moles of water are produced when 96.0 g of hydrazine (N_2H_4) are burned in a rocket.

The balanced equation for this reaction is:

$$N_2H_4 + O_2 \rightarrow 2H_2O + N_2$$

The molar mass of hydrazine is 32.0 g. In this problem 96.0 g of N_2H_4 are burned. How many moles of N_2H_4 is this?

$$\frac{96.0\ g}{32.0\ g/mol} = 3.0 \text{ mol of } N_2H_4$$

The balanced equation tells us that

$$1 \text{ mol } N_2H_4 \rightarrow 2 \text{ mol } H_2O$$

This must mean that

$$3 \text{ mol } N_2H_4 \rightarrow 6 \text{ mol } H_2O$$

We see that 96.0 g or 3 mol of hydrazine produce 6 mol of water. Since the molar mass of water is 18.0 g, then 6 × 18.0 or 108 g of water would be formed.

You may use ratio and proportion to calculate moles of H_2O. This procedure may be handy when dealing with complex numbers.

$$\frac{1 \text{ mol } N_2H_4 \text{ (from equation)}}{3 \text{ mol } N_2H_4 \text{ (given)}} = \frac{2 \text{ mol } H_2O \text{ (from equation)}}{X \text{ mol } H_2O \text{ (required)}}$$

Using numbers only we have

$\frac{1}{3} = \frac{2}{X}$ By cross multiplying

we obtain $X = 6$; $\therefore$ 6 mol of H_2O are produced.

<u>Example 2</u>

When iron rusts it combines with oxygen in the air to form iron oxide, Fe_2O_3, according to the following balanced equation:

$$3O_2 + 4Fe \rightarrow 2Fe_2O_3$$

Suppose we wish to calculate the mass of oxygen required for the complete reaction of 83.7 g of iron.

The molar mass of iron is 55.8 g. In this problem 83.7 g of iron react. How many moles of iron react?

$$\frac{83.7 \text{ g}}{55.8 \text{ g/mol}} = 1.50 \text{ mol of iron}$$

The balanced equation tells us that

$$\underline{3} \text{ mol } O_2 \text{ reacts with } \underline{4} \text{ mol Fe}$$

This must mean that

$$3/4 \times 1.50 \text{ or } 1.12 \text{ mol } O_2 \text{ react with } 1.50 \text{ mol Fe}$$

To calculate moles of O_2 required, you may use ratio and proportion if you find it easier.

$$\frac{3 \text{ mol } O_2}{X \text{ mol } O_2} = \frac{4 \text{ mol Fe}}{1.50 \text{ mol Fe}}$$

$$4X = 4.50$$

$$X = 1.12 \text{ mol of } O_2 \quad \text{(use significant figures)}$$

Since the molar mass of oxygen is 32.0 g, the mass of 1.12 mol of oxygen is

$$1.12 \text{ mol} \times \frac{32.0 \text{ g}}{\text{mol}} = 35.8 \text{ g}$$

Therefore you may conclude that 35.8 g of oxygen will be required to react with 83.7 g of iron.

Exercise 1

1. Methane, the main constituent of natural gas, has the formula CH_4. The products of combustion are carbon dioxide and water. Use spaces provided to answer questions (a) to (d).

(a) Write the balanced equation for burning methane.

(b) One mole of methane produces how many moles of water vapor? (Show work.)

(c) One-eighth mole of methane would produce how many moles of carbon dioxide?

(d) How many moles of water vapor would be produced by 4.0 g of methane?

If your answer to 1(d) was not "$\frac{1}{2}$", do the question below.

2. Unbalanced: $CH_4 + O_2 \rightarrow CO_2 + H_2O$

Fill in the blanks with the correct coefficients.

(a) Balanced: ___ CH_4 + ___O_2 → ___CO_2 + ___H_2O

(b) Mole ratio: ____mol + ____mol → ___mol + ____mol

(c) 1/8 mol CH_4 → $\underline{X}$ mol CO_2

X = ___ mol of CO_2

Therefore 1/8 mol CH_4 produces ___ mol of CO_2.

(d) molar mass of CH_4 = 16 g

Therefore 4 g of CH_4 = $\frac{\underline{\quad}\text{g}\underline{\quad}}{\text{g/mol}} = \frac{4.0\text{ g}}{\underline{\quad}\text{g/mol}}$ = 0.25 mol

$\underline{1}$ mol of CH_4 produces ___ mol of water vapour (see equation (a))

0.25 mole CH_4 (4 g) produces ___ mol water vapour.

To do: Practice Exercise - pages 221-222; problems 4 (a-d), 5 (a-d), 8 (a-b), 9 (a-c).

Check your answers with the textbook. If you are unable to solve some problems, please indicate which particular problem you would like explained.

Do not give up too soon. Follow examples in this lesson and your textbook.

Exercise 2

1. Given: $CaCO_3 \rightarrow CaO + CO_2$.

(a) If 150.1 g of $CaCO_3$ are used in the above reaction, how many moles is this?

(b) How many moles of CO_2 would be produced by heating 150.1 g of $CaCO_3$?

(c) How many grams of CO_2 would be produced?

2. Given: $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$

   (a) If 4.0 g of hydrogen are produced by the above reaction, how many moles of hydrogen molecules are produced?

   (b) How many grams of calcium would be required to produce 4.0 g of hydrogen?

3. Given that 48.6 g of magnesium, Mg, react with oxygen to yield a white solid, MgO:

   (a) Write the balanced equation for this reaction.

   (b) Calculate the number of moles of MgO that will be produced.

4. What mass of Cl_2 is needed to produce 63.5 g of I_2 when chlorine reacts with NaI according to the following equation (unbalanced):

$$NaI + Cl_2 \rightarrow NaCl + I_2$$

5. Given: $Pb(NO_3)_2 + NaI \rightarrow NaNO_3 + PbI_{2\,(s)}$

The above reaction occurs in your Exp. 3, Lesson 5. What mass of yellow precipitate, PbI_2, will be produced when 33.1 g of $Pb(NO_3)_2$ react. Show all steps below:

6. (a) If 71.0 g of chlorine gas and 10.0 g of hydrogen gas are allowed to react in a reaction chamber, calculate the mass of hydrogen chloride gas (HCl) produced.

(b) Which gas, if any, is left in excess, and state the amount.

(Appendix 10A may be helpful for the above section.)

F. Types of Chemical Reactions

To do: Please read Section 8-10, pages 218-220 of your textbook. Further examples of different types of reactions appear below.

(1) Synthesis Reaction

Refer to Experiment 7. You will note that iron combined with sulphur upon heating to produce an entirely new product, iron sulfide.

The equation for the above reaction is:

$$Fe + S \rightarrow FeS$$

Another example, is synthesis of ammonia, which is of great industrial use. Here, nitrogen reacts with hydrogen, under special conditions to produce ammonia (NH_3). The equation for this reaction is:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Exercise 3 : If 7.0 kg of nitrogen react, calculate the mass of hydrogen required for this reaction.

(2) Decomposition Reaction

You may have experienced a demonstration in other science courses where oxygen gas was produced by heating a compound such as potassium chlorate. This reaction shows decomposition of a compound into other simpler substances as shown by the equation below:

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

Probably you have also seen a demonstration on electrolysis of water, where the compound water decomposes into its elements as shown by the following equation:

$$2H_2O \rightarrow 2H_2 + O_2$$

(3) Rearrangement or Replacement Reaction

In this type, two compounds react in such a way as to produce two new compounds. This reaction was evident in Experiment 7 when hydrochloric acid reacted with iron sulfide to produce an obnoxious smelling gas and another chemical residue. This reaction may be represented by the following equation:

$$FeS + 2HCl \rightarrow H_2S + FeCl_2$$

(4) Metal-ion Reaction

Refer to Experiment 5. You will note that magnesium ribbon reacted with acetic acid to produce hydrogen gas and magnesium acetate. Here, magnesium metal reacted with the hydrogen ion present in acetic acid to produce the new compounds. The equation for this reaction is:

$$Mg + 2HCH_3COO \rightarrow H_2 + Mg(CH_3COO)_2$$

Note that HCH_3COO yields H^+ ions and CH_3COO^- ions when in solution.

These ions are not represented as such in this course for the sake of simplicity.

(5) Ion-ion Reaction

Your second experiment in this course involved interaction of ions to produce a new compound such as the yellow precipitate you observed when lead nitrate solution reacted with potassium iodide solution. The yellow precipitate is a result of the reaction between the lead ions, Pb^{+2}, and the iodide ions, I^-, to produce lead iodide, PbI_2, a yellow precipitate.

Although the molecular equation does not show this, it is worthwhile to keep in mind the reacting ions which produce a new substance. The over all equation for the above reaction is:

$$Pb(NO_3)_2 + 2NaI \rightarrow PbI_2 + 2NaNO_3$$

Note: Writing of equations using ions will be dealt with in your next chemistry course.

To do: Practice Exercise, Page 222, Problem 7.
Self-Test, Page 222, Problem 4.

Please check your answers with the textbook.

N.B. So far in this course, you have learned what an ion is, but since chapter 7 will be covered in Chemistry 20, you have no way of knowing which compound is ionic, and which is molecular, so some of your answers may have been wrong. In the following exercise, assume molecules are involved unless ions are clearly indicated. Remember, if something does not have a (+) or (−) charge, it is not an ion.

Exercise 4 (Send for correction)

Classify the following reactions as to type. Use the space provided.

(a) $S_8 + 8O_2 \rightarrow 8SO_2$ ________________

(b) $Al_2(SO_4)_3 + 3Ba(OH)_2 \rightarrow 3Al(OH)_3 + BaSO_4$ ________________

(c) $Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 + 2NaCl$

$2Na^+ + SO_4^{-2} + Ba^{+2} + 2Cl^- \rightarrow BaSO_{4(s)} + 2Na^+ + 2Cl^-$ ________________

(d) $Mg + ZnSO_4 \rightarrow Zn + MgSO_4$

$Mg + Zn^{+2} + SO_4^{-2} \rightarrow Zn + Mg^{+2} + SO_4^{-2}$ ________________

(e) $H_2S \rightarrow H_2 + S$ ________________

(f) $Ca + Cl_2 \rightarrow CaCl_2$ ________________

(g) $N_2 + 6Li \rightarrow 2Li_3N$ ________________

(h) $2K + 2H_2O \rightarrow H_2 + 2KOH$ ________________

G. Experiment 10: A QUANTITATIVE STUDY OF A CHEMICAL REACTION

This experiment is outlined in your Laboratory Manual on pages 97-99. We will have to modify it to some extent, since you will depend on the material found in the laboratory kit.

Materials:

* sodium iodide (NaI)
* lead nitrate $Pb(NO_3)_2$
* 1 - eye dropper
* 5 test tubes (13 × 100 mm)
tap water
10 mL graduated cylinder
balance
* 100 mL beaker
labelling tape
test tube stand (a glass will work)
a ruler (marked in cm)
* glass stirring rod

Purpose:

In this experiment, your laboratory experiences will require careful attention to the amounts of chemicals used, and close observation of the results. The mole concept will be emphasized. You will see that for any given reaction, one of the two materials is often in excess. How much product is produced will depend on the substance which is present in the lesser amount.

Procedure:

Put five 13 × 100 mm test tubes in a test tube support (see diagram page 11). Label them 1, 2, 3, 4, 5 (Use masking tape or scotch tape.) Secure labels near the top of the test tubes.

Weigh exactly 3.0 g of sodium iodide (NaI). Caution: Read Rule 2 page 8 of the Laboratory Manual before you start the experiment. If using paper to protect the pan, weigh it first, then add 3 g of the compound.

Place the sodium iodide into a clean, dry, beaker, and add enough water to make 40 cm^3 of solution. Use your graduated cylinder to measure out the water. Stir thoroughly until the compound dissolves completely.

Pour 4.0 cm^3 of the sodium iodide solution into each test tube. Save the remaining solution for later use. Next weigh exactly 3.3 g of lead nitrate $Pb(NO_3)_2$. Observe caution as before. Remove and store the previous contents from the beaker, wash and dry it. Place the lead nitrate $Pb(NO_3)_2$, into the clean beaker. Add enough water to make 20 cm^3 of solution. Stir until the compound dissolves completely. Make sure the stirring rod is clean. (Rinse before using.) See * below.

Using your graduated cylinder (make sure it is clean and dry) add the following amounts of lead nitrate solution to each test tube. Save the remaining solution for later use.

Tube	Sodium Iodide	Lead Nitrate
1	4 cm^3	0.5 cm^3
2	4 cm^3	1.0 cm^3
3	4 cm^3	2.0 cm^3
4	4 cm^3	3.0 cm^3
5	4 cm^3	4.0 cm^3

Place a piece of plastic over each test tube and shake to mix the solution thoroughly. (Be careful not to spill any solution.) Place the five test tubes in a glass or some other holder so that the test tubes rest in an upright position so that the precipitates will settle evenly. This is important in order to measure their heights accurately. Let the test tube stand for at least two hours. When the precipitates have settled measure and record the height of each to the nearest tenth of a centimetre.

*4 cm^3 of each solution should have 0.002 moles of each compound now.

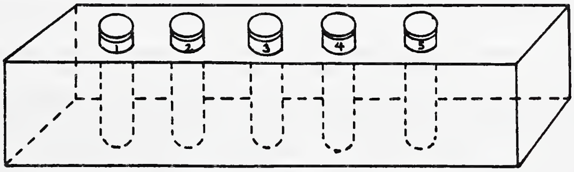
Record:

Test Tube #	Height of Precipitate (in mm)	Data
1	__________	6 mm
2	__________	12 mm
3	__________	24 mm
4	__________	24 mm
5	__________	24 mm

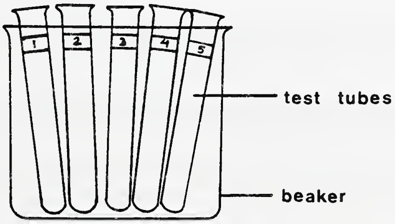
(Note: Heights in Test Tube 3, 4, and 5 should be almost equal. If not, use data given above to answer questions on pages 12 to 15)

A Test Tube Support:

I Use a cardboard box whose width is less than the length of a test tube. Make five holes on one side of box, for the test tubes to fit in snugly. Set the box on its side and insert the test tubes as required.



II You may also use a glass beaker to hold the test tubes. However, the test tubes should be labelled first.



To Label Your Test Tubes:

I You may use masking tape or

II Write the label, (test tube number) on a strip of paper and scotch tape it near the top of the test tube.

Note: The strength of each solution is identical. This means that 1 cm^3 of lead nitrate solution has as many moles of solution as 1 cm^3 of sodium iodide solution. The molecular weight of $Pb(NO_3)_2$ is 331. The molecular weight of NaI is 150. For this reason, a given volume will not have identical numbers of grams of solute, even though the number of moles is the same.

Questions on Experiment 8:

1. Predict what adding one more drop of lead nitrate to each of the test tubes will do. Will it change the height of the precipitate deposit in any of them? Record your predictions in the table below.

Test tube	Predicted height (increased, decreased, no change)
#1	
#2	
#3	
#4	
#5	

2. Add one drop of lead nitrate to each test tube. Observe very closely to see if any precipitate forms. Record your observation in table below.

Test tube	Observed height (increased, decreased, no change)
#1	
#2	
#3	
#4	
#5	

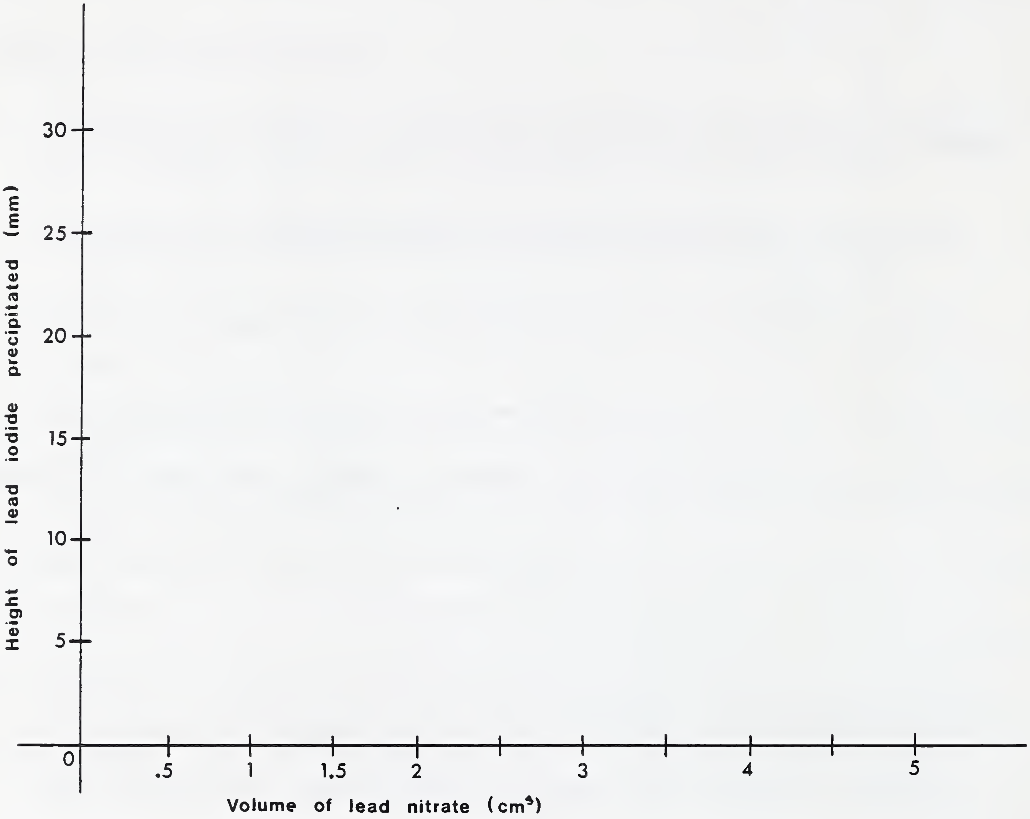
3. Add one drop of sodium iodide to each test tube. Observe very closely to see if any precipitate forms. Record your observations in table below.

Test tube	Observed height (increased, decreased, no change)
#1	
#2	
#3	
#4	
#5	

4. From your observations in questions 2 and 3 state your conclusions.

5. If the strength of solutions and measurements were exact, would you expect more precipitate to form in test tube #3 upon adding one drop of lead nitrate or one drop of sodium iodide? Explain.

6. Graph the height of the precipitates as a function of the volume of lead nitrate. Connect the points with a line.



7. What is the ratio of the volume of sodium iodide to the volume of lead nitrate for a complete reaction?

____

____

8. Write the <u>word</u> equation for the reaction between sodium iodide and lead nitrate.

____

____

9. If the charge on the lead ion is $^{+2}$ and the charge on the iodide ion is $^{-1}$, write the correct formula for the compound, lead iodide.

10. Propose a hypothesis to explain the ratio in your answer to question 7.

11. (a) In which test tube(s) were there apparently lead ions that did not react?

    (b) Give a reason for your answer.

12. Using all the information obtained in this experiment, how many moles of sodium iodide do you think are required to react completely with one mole of lead nitrate?

13. Write a balanced chemical equation for the reaction between sodium iodide and lead nitrate.

14. How many moles are there in 7.5 kg of NaI? (Hint: 1 kg:1.0×10^3 g) (1 mol of NaI = 150 g)

PLEASE NOTE: If you borrowed a beam balance, set of weights, a thermometer and a graduated cylinder for use in your course, this was the last lesson that required its usage. Please return it to the Alberta Correspondence School, if you have not already done so.

END OF LESSON 8

LESSON RECORD FORM

1240 Chemistry 10

Revised 90/06

FOR STUDENT USE ONLY

Date Lesson Submitted

____

Time Spent on Lesson

____

(If label is missing
or incorrect)

File Number

____

Lesson Number ____________

FOR SCHOOL USE ONLY

Assigned
Teacher: ________

Lesson Grading: ________

Additional Grading
E/R/P Code: ________

Mark: ________

Graded by: ________

Assignment Code: ______

Date Lesson Received:

____

Lesson Recorded ______

Student's Questions and Comments

Apply Lesson Label Here

Name

Address

Postal Code

Please verify that preprinted label is for correct course and lesson.

Teacher's Comments:

____

Correspondence Teacher

St. Serv. 21-89

ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

(1) All pages are numbered and in order, and no paper clips or staples are used.
(2) All exercises are completed. If not, explain why.
(3)· Your work has been re-read to ensure accuracy in spelling and lesson details.
(4) The Lesson Record Form is filled out and the correct lesson label is attached.
(5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

First Class

Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope. Correspondence lessons will travel faster if first-class postage is used.**

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

METALS

METALS IN GENERAL

A. History of Metals
B. Metallurgy in the Future
C. Sources of Metals
D. Prices of Metals
E. Production of Metals
F. Definition of a Metal
G. States of a Metal
    (1) Crystal State
    (2) "Plastic" State
    (3) Liquid State
H. Metals on the Periodic Table
I. Copper - From Beginning to End

A. History of Metals

Civilization has progressed from the Stone Age to the Bronze Age to the Iron Age, depending on the material man used to make his weapons. The reason this is a progression, is that iron weapons were much stronger than bronze weapons, which were in turn stronger than stone weapons. In each broad age group, many important things about each metal were discovered. There are no clear-cut dates when one age ended, and when another began, since there was much overlap, and different civilizations progressed faster than others.

Metals have been known since very early times in man's history. Probably the first metals known to man were copper, silver, and gold, since these metals can be found in the pure state, and uncombined with other elements. Man soon found good uses for these metals, and the search began for other metals. In time, man could reduce ores of metals, and thereby obtain the pure metal from the compound. As a result, tin, lead, iron, and mercury were also known in early times. With the metals he had, man started producing various alloys by melting various metals together.

The Bronze Age was born when it was discovered that copper and tin could be melted together to form an alloy which was harder and more durable than either metal alone. This age was followed by the Iron Age.

The Iron Age was divided into three periods. In the first period, iron could not be heated high enough to melt it, and all of the work on iron had to be done by heating the iron as high as possible, and then pounding it by hand with heavy hammers. In the second period, which began in the fourteenth century, iron makers could melt the iron, and had more control over the quality of iron which was produced. The third period, which began in the middle of the nineteenth century, saw the use of the Bessemer process for making steel. This last period is also called the Age of Iron and Steel.

Of course, in the various periods of the Iron Age, man progressed in his knowledge of other metals. Besides learning many basic processes in the treating of iron, and various refinements thereof, man discovered how to treat other metals in various ways, such as the smelting of zinc, copper, and lead. Later discoveries included electrolytic copper and aluminum refining, as well as the cyanide process for the extraction of gold.

B. Metallurgy in the Future

Early man discovered metals one by one, and today, man has discovered all naturally occuring elements and metals. Any new metals which will be discovered, will be radioactive metals. These will be produced in the laboratory, and will not be too useful commercially.

Although the number of metals may be limited, just as the number of letters in the alphabet may be limited, there is no limit to the various ways in which metals can be combined to produce different alloys. Besides searching for new alloys, scientists are also searching for new and better ways of producing alloys, along with better ways of treating metals and alloys to give the most desirable properties. Something that works well for one metal or alloy may not work at all for the next metal or alloy since different metals have great differences in their physical, and to some extent, chemical properties.

In the extractive metallurgy field, many of the best mining sites for particular metals may be mined out, and instead of working with an ore which has eight per cent copper, one may have to work with an ore which has one per cent copper. In order for the mine to be economical, entirely different methods might be required to extract the metal from the ore.

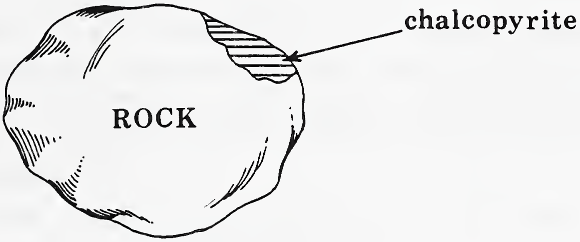
C. Sources of Metals

By far, the most important source of metals is the earth's crust. However, one rarely finds a pure metal. Almost always, we find ores which contain important metals, and these ores are then mined and milled. Iron is one of our most important metals. Some reasons why iron is so important are that iron ores are easy to reduce to the metal, it is very cheap, it is found in great quantities, it is a very useful structural material, and it is easily worked into various shapes.

There may be little resemblance between the ore from which a certain metal is obtained, and the color and properties of the metal itself. An ore of a given metal may have oxygen, sulphur, or carbonates, or a number of other elements or compounds which are chemically combined with the metal. These other elements give the ore different properties from the metal itself. For instance, hematite, (Fe_2O_3), which is the most important ore mineral of iron, is usually red in color, and it is not magnetic. Iron is gray and magnetic.

Most ores are generally found in certain regions of the world, and they are usually not scattered all over. However, ores of certain metals may be very abundant in one country, but the ore itself may be treated in another country. This is the case with aluminum since much electric power is needed to refine aluminum. See Section E for the important countries for certain common metals.

It should also be noted that high grade sources of many metals are rapidly being used up, and it is important to find economical means of treating lower grades of ore. Today, scrap metal is being recycled, and it is becoming an important source of metals. What a low grade ore is, depends on the metal. For instance, it may not be economical to mine in an area which has 15 - 20% iron, although it may be economical to mine in an area which has one per cent copper or less. The diagram illustrates what a low grade ore is. In a mill, the whole rock may be crushed and ground and treated in a mill, only to get the little chunk of chalcopyrite. This chalcopyrite is only about one-third copper, which must then be sent to the smelter to obtain pure copper.



D. Prices of Metals

There are a number of different factors which influence the price of a metal. Some of these factors are: 1. how badly the metal is needed, 2. how much of the metal is available, 3. how easily it is mined and refined, 4. and the expense of forming operations.

Regarding the first factor, the price of a metal can often be increased by finding new uses for the metal. However, it would also help if there wasn't a cheaper material available which could do the same job. Let us assume that silver and copper are excellent conductors of electricity, and that both would do a certain job equally well. Naturally, copper would be chosen as it is much cheaper. For some applications, you may need several specific qualities in a metal, and a metal which has a number of good qualities would be worth more than a metal with just one or two good qualities.

The availability of a metal greatly influences its price. The huge deposits of iron ore which are found in many parts of the world has a lot to do with the relative cheapness of iron and steel. If silver were as plentiful as iron, it would cost much less than it does today.

How easy it is to obtain the metal from the ore also greatly influences the price. Aluminum would illustrate this point well. Although aluminum ores are very plentiful, it used to be extremely difficult to obtain pure aluminum from its ore. Before Hall discovered a new process for extracting aluminum, aluminum cost eight dollars a pound, however, after his discovery, the price dropped to fifteen cents a pound. Even with the new process, huge amounts of electricity are required, and scientists are continually searching for cheaper methods of extracting aluminum. The purity of the metal required and the processes used to obtain this purity also influence the price of a metal. For many applications, the odd extra atom may not make much difference, but the conductivity of copper, for instance, can be decreased by a half, with only 0.3% arsenic.

The forming operations a metal requires also influences its price. When a certain metal is bought, one usually does not want to use a chunk of the metal, but one wants to be able to shape and form it in a certain manner. Stainless steel, for instance, is very hard, and it is much harder to form it into a certain shape than lead. The tools used to shape stainless steel would have to be harder, and would wear out faster than those used for lead. If everything else were the same, the greater expense for the forming operation of stainless steel would make it more expensive.

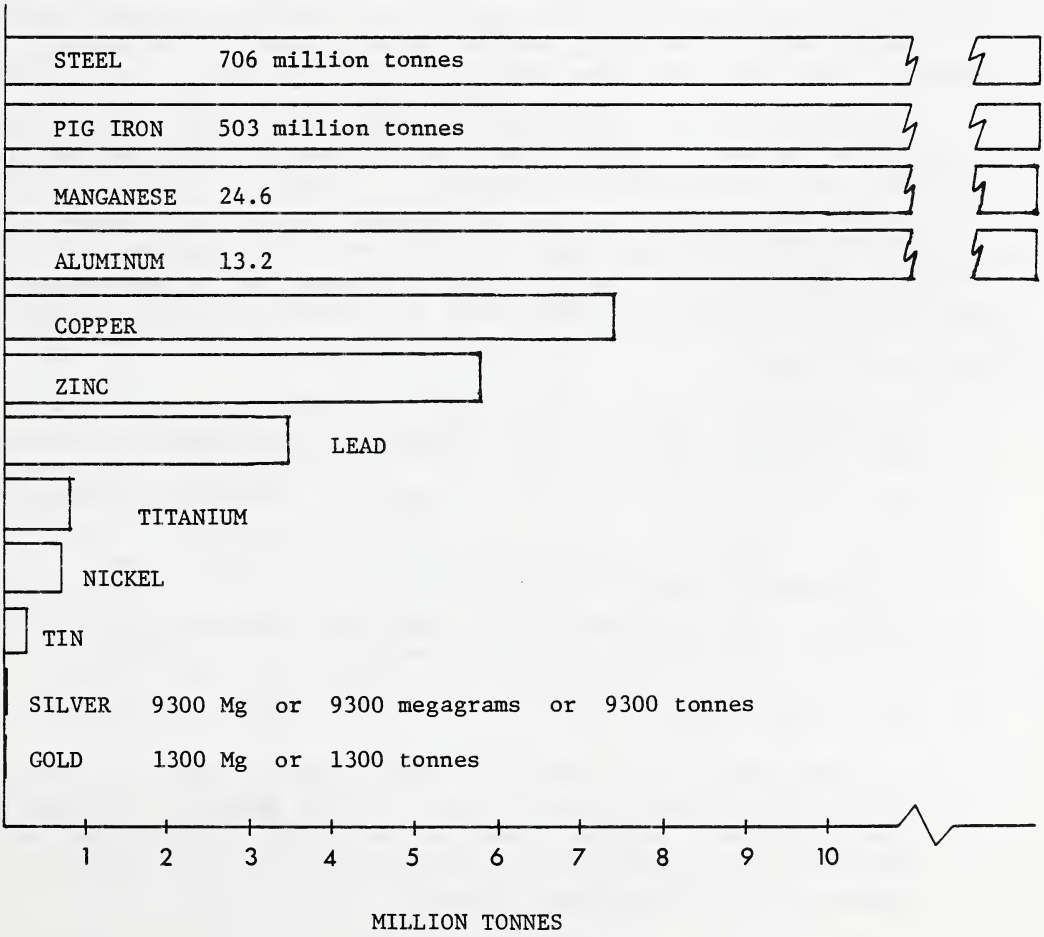
Below, some recent prices of common metals will be given. The prices change very quickly, and therefore the price at any time would be slightly different than the prices below, however, this is just meant to be a rough guide for the relative value of various metals.

		1976 prices
Gold	$4.20/g	($130.00/ounce)
Silver	$0.14/g	($ 4.40/ounce)
Tin	$9.00/kg	($ 4.07/pound)
Nickel	$5.30/kg	($ 2.41/pound)
Copper	$1.43/kg	($ 0.65/pound)
Aluminum	$1.06/kg	($ 0.48/pound)
Zinc	$0.83/kg	($ 0.38/pound)
Lead	$0.58/kg	($ 0.27/pound)
Manganese	$0.41/kg	($410.00/tonne)
Steel (carbon)	$0.23/kg	($230.00/tonne)
Pig Iron	$0.06/kg	($ 60.00/tonne)

E. Production of Metals

The following graph will give the approximate tonnages of metals produced in the world in 1974. Different countries produce varying amounts of the different metals, but below it will be indicated which countries are leading producers of some of the more common metals. The United States is noted for its production of steel, silver, lead, copper, and aluminum. The U.S.S.R. is noted for iron and steel and manganese. Canada is an important producer of nickel, silver, and zinc. South Africa is noted for its gold, Malaysia for its tin, and Australia for its titanium. Other important producers of these and other metals can be found in a yearbook.

APPROXIMATE METAL PRODUCTION IN 1974



F. Definition of a Metal

At first glance, this may seem like a trivial point, since we all know that most cutlery, as well as cars and trains consist of metals, and that things like paper and water are not metals. However, what distinguishes a metal from a nonmetal? Many metals we come in contact with are hard, however, does this imply that once some copper or iron has melted, it is no longer a metal? As we will see later in this course, there are a large number of metals, and a number of these have properties which are not related to properties of other metals. A simple definition of a metal, which is a property of all metals, and which, at the same time, is not a characteristic property of nonmetals, is that a metal is an element which readily loses electrons. As a result, metals, are good conductors of heat and electricity, and they are opaque and lustrous. Also, metals are often malleable, ductile, and generally heavier than other elemental substances. However, some metals have more of a certain property than other metals, and some nonmetals also have some of the above properties, so one has to be careful before something can be called a metal. For example, wood is opaque, since you can't see through it, although wood is not a metal.

To illustrate some differences between metals, the melting point of mercury is so low, (-40°C), that it is a liquid at room temperature. In contrast, the melting point of tungsten is 3410°C. Lithium is so light that it can float on water, yet osmium is so heavy that a milk bottle full of osmium would weigh 22 kg. Lithium is as soft as wax and other metals like iron are extremely hard. Despite these great differences in their physical properties, all metals react in a similar manner chemically.

G. States of a Metal

Unfortunately, this is a very difficult section since it deals with positions of atoms which are too small to be seen, and which one generally does not think of when one looks at anything. Although it would be beyond the scope of this course to describe the exact positions of the atoms in the various metals, hopefully the simplified descriptions that are given, will be adequate to explain exactly what is done when a metal is heated and worked. Three states of metals will be discussed, which will be referred to as the crystal state, the plastic state, and the liquid state.

(1) The Crystal State

The atoms in a metal are in their crystal state at a temperature well below the metal's melting point. The atoms are arranged in an orderly manner, and the atoms do not readily readjust themselves when the metal is deformed in some way. An iron bar will be illustrated in the following manner.

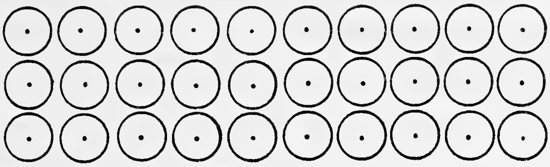


Figure 2: Metal in the crystal or plastic state

The circles represent the outer influence of the electrons. The dots represent the nucleus. Either method may be used in the following diagrams, depending on which illustrates the point in the best and easiest way.

Now, when the bar is bent, while the atoms are in the crystal state, the bar may look as follows.

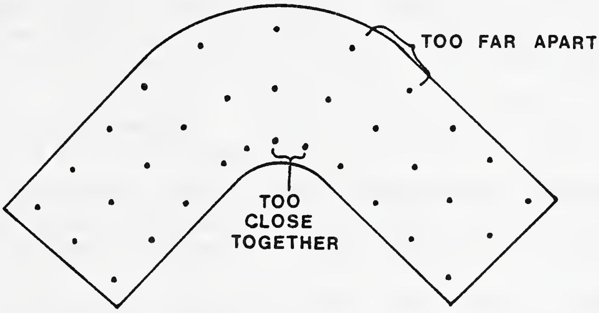


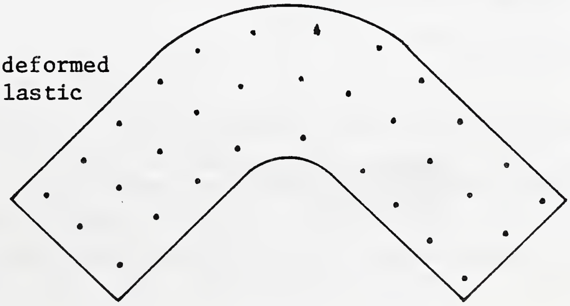
Figure 3: A metal deformed in the crystal state

The middle row of atoms are happy the way they are, since they are as far as they would like to be from the neighboring atoms. However, in the bottom row, the atoms are squeezed together more than they would like to be, and on top, the atoms are further apart than they would like to be. In the crystal state, the atoms have a difficult time readjusting themselves into a comfortable position, and in the crystal state, the metal strip will break if you bend it too often, or too much.

(2) The Plastic State

When a metal is in the plastic state, it is still a solid, and a piece of metal will have the same shape in the plastic state as it did in the crystal state. However, if one deforms the metal in some way, the individual atoms will readjust themselves so that there is no stress on them. The iron bar of figure 2 will look as follows if it were bent in the plastic state.

Figure 4: A metal deformed in the plastic state



In the figure, the middle atoms are happy as before. However, the other atoms have also somehow readjusted, and now all atoms in the bottom row and in the top row are as far from the next atom as they would like to be. After this metal is cooled, the bent bar would be just as ductile as the metal originally was when it was in the crystal state. However, if the metal were just bent in the crystal state, without being heated, it would be more brittle after it was bent, than it was before.

The temperature that one needs, in order to get the metal into the plastic state, depends on the melting point of the metal. The higher the melting point, the higher is the temperature that one needs to get the metal into the plastic state. The melting point of some metals, such as tin and lead, is so low that the atoms are in the plastic state at room temperature. Even if atoms are in the plastic state already, heating the metal to a still higher temperature would make the metal atoms readjust themselves faster. However, to be in the solid plastic state, the temperature must be higher than in the crystal state, but it must be lower than in the liquid state. Iron, for instance, is in the crystal state below 450°C. Between 450°C and 1540°C, iron is in the plastic state, and above 1540°C, iron is a liquid.

Perhaps an analogy would illustrate the difference between the crystal and plastic state. Suppose that one had 100 people who talked five different languages. Also suppose that 100 people were put into a room just big enough for all people, and that no one new the language that the next person spoke. Assume that the feet were glued to the floor, and that no one could turn after he got inside the room. Once the people start talking, one person may realize that a person who is talking in his own language is just behind him or just past his neighbor. Unfortunately, since his feet are glued to the ground, he can't do anything about it and therefore feels very uncomfortable. This is exactly how a metal atom feels if it is bent in the crystal state. Now let us assume that each person may take one step in any direction. Everyone would be very happy then, as all would be beside one or more people who talk his own language. This is how a metal atom feels if it is bent in the plastic state. Looking from the outside, you wouldn't see any difference. The room is still crowded with people, and the cup, or whatever shape the metal had, would still look like the same cup, however, the internal changes makes the cup much more ductile than before.

(3) The Liquid State

When something is in the crystal state, or plastic state, it would still have the same shape. The atoms in these two states would be in a definite pattern and each atom will be a certain distance from the next atom. In the liquid state, the solid loses its structure, just like an ice statue would, after it had melted. The atoms would not be in a definite pattern, and they could slide over each other readily.

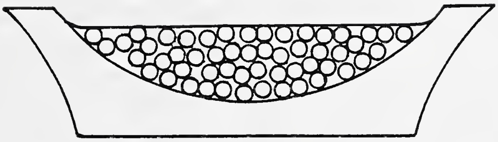


Figure 5: A liquid metal

In the above diagram, note that the atoms are varying distances apart, as contrasted with Figure 2.

N.B. Please keep in mind that the above discussion, particularly on the plastic state, is extremely simplified, and that metals always extend in three dimensions and not two.

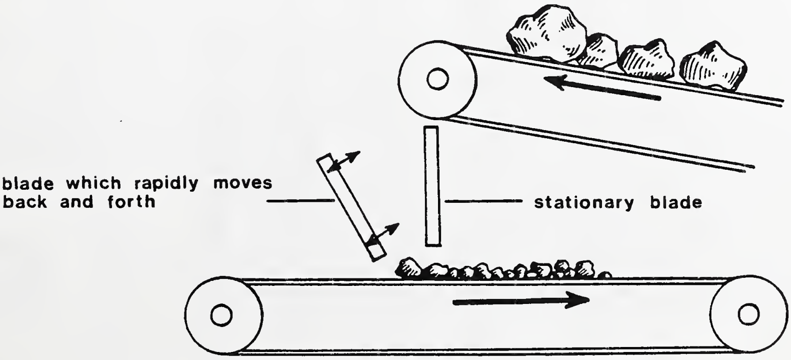
H. Metals on the Periodic Table

Attached is a periodic table with all elements present. As can be seen, four fifths of all elements are metals, since all elements with lines through them are metals.

I. Copper - From Beginning to End

Earlier, we mentioned that an ore may only have a very small amount of copper in it. Below, we will discuss in greater detail how an ore with perhaps only one per cent copper is treated to get 100% copper. Very similar processes that apply to copper also apply to many other metals as well.

First of all, the ore body is broken up into large rocks in the mine. The rocks, which could be any size up to 1 m or 1.5 m long, are sent through various types of crushers, usually by means of conveyor belts. Larger rocks enter the crushers from the top and smaller rocks fall on to the conveyor belt below the various crushers.



After a series of crushing and grinding operations, the large rock is broken up into dustlike particles which are then mixed with water. Various chemicals are put into the water, which are greatly attracted to ores, but which have no effect on the unwanted rock which is called the gangue. These added chemicals are not very soluble in water, and when air is blown into the water, the chemical, along with the ore, floats to the top, leaving the gangue behind. This process is called "flotation."

By this process, you should get most of the ore out, but not all of it, since some of the dustlike particles will still have both rocks and minerals on it.

At this point, the ore is roasted, which means to heat the ore in air, since it is easier to obtain the pure metal from an oxide than the carbonate or sulfide. If your ore is Cu_2S, the equation for the roasting of this ore is

$$Cu_2S + O_2 \rightarrow Cu_2O + SO_2\uparrow$$

At this point, the oxide may be combined with coke to get copper.

$$Cu_2O + C \rightarrow Cu + CO\uparrow$$

The previous equation implies that pure copper is obtained, however all kinds of impurities may still be present in the "copper". Although this "copper" is 99% pure, it may still have minute amounts of silver, gold, zinc, and other metals. This copper is then made more pure by a process called "electrolysis" in which an electric current is used to get very pure copper. In electrolysis, a current is passed through the impure copper, resulting in copper II ions. These ions are then collected at the other end. It is necessary to get extremely pure copper since a small amount of impurity will have a large effect on the conducting properties of copper.

Periodic Table of Elements

KEY

atomic number	1	
atomic mass (g/mol) ( ) most stable isotope	1.01	
melting point (°C) (w-white)	-259	
boiling point	-253	
density (g/cm^3) (if gas at S.T.P. then density is for liquid at boiling point)	0.071	
name of element	hydrogen	
symbol of element (slant: synthetically prepared)	H	
electronegativity	2.1	
ionic charges (most common second most common)	1+ 1-	
number of atoms per molecule of most common type	2	

▨ Metals

☐ Non-metals

GROUP IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1 H 1.01 2.1 -259 1+ -253 1- 0.071 2 hydrogen																	2 He 4.00 - -272 -269 0.126 helium
3 Li 6.94 1.0 180 1+ 1347 0.53 lithium	4 Be 9.01 1.5 1278 2+ 2970 1.85 beryllium											5 B 10.81 2.0 2079 2550 2.34 boron	6 C 12.01 2.5 3550 4827 2.26 carbon	7 N 14.01 3.0 -210 3- -196 0.81 2 nitrogen	8 O 16.00 3.5 -218 2- -183 1.14 2 oxygen	9 F 19.00 4.0 -220 1- -188 1.505 2 fluorine	10 Ne 20.18 - -249 -246 1.20 neon
11 Na 22.99 0.9 97.8 1+ 883 0.97 sodium	12 Mg 24.31 1.2 649 2+ 1090 1.74 magnesium											13 Al 26.98 1.5 660 3+ 2467 2.70 aluminum	14 Si 28.09 1.8 1410 2355 2.33 silicon	15 P 30.97 2.1 44.1 w 3- 280 w 1.82 w 4 phosphorus	16 S 32.06 2.5 113 2- 445 2.07 8 sulfur	17 Cl 35.45 3.0 -101 1- -35 1.56 2 chlorine	18 Ar 39.95 - -189 -186 1.40 argon
19 K 39.10 0.8 63.7 1+ 774 0.86 potassium	20 Ca 40.08 1.0 839 2+ 1484 1.55 calcium	21 Sc 44.96 1.3 1541 3+ 2831 3.0 scandium	22 Ti 47.90 1.5 1660 4+ 3287 3+ 4.51 titanium	23 V 50.94 1.6 1890 5+ 3380 4+ 6.1 vanadium	24 Cr 52.00 1.6 1657 3+ 2672 2+ 7.19 chromium	25 Mn 54.94 1.5 1245 2+ 1962 4+ 7.43 manganese	26 Fe 55.85 1.8 1535 3+ 2750 2+ 7.87 iron	27 Co 58.93 1.8 1495 2+ 2870 3+ 8.9 cobalt	28 Ni 58.71 1.8 1453 2+ 2732 3+ 8.9 nickel	29 Cu 63.55 1.9 1083 2+ 2567 1+ 8.96 copper	30 Zn 65.38 1.6 420 2+ 907 7.13 zinc	31 Ga 69.74 1.6 29.8 3+ 2403 5.90 gallium	32 Ge 72.59 1.8 937 4+ 2830 5.32 germanium	33 As 74.92 2.0 - 3- 615 5.73 arsenic	34 Se 78.96 2.4 217 2- 685 4.79 selenium	35 Br 79.90 2.8 -7.2 1- 58 3.12 2 bromine	36 Kr 83.80 - -157 -152 2.41 krypton
37 Rb 85.47 0.8 38.9 1+ 688 1.53 rubidium	38 Sr 87.62 1.0 769 2+ 1384 2.5 strontium	39 Y 88.91 1.3 1522 3338 4.47 yttrium	40 Zr 91.22 1.4 1852 4377 6.51 zirconium	41 Nb 92.91 1.6 2468 4742 8.6 niobium	42 Mo 95.94 1.8 2617 4612 10.2 molybdenum	43 *Tc* 98.91 1.9 2172 4877 11.5 technetium	44 Ru 101.07 2.2 2310 3900 12.4 ruthenium	45 Rh 102.91 2.2 1966 3727 12.4 rhodium	46 Pd 106.40 2.2 1552 2+ 3140 4+ 12.0 palladium	47 Ag 107.87 1.9 962 1+ 2212 10.5 silver	48 Cd 112.41 1.7 321 2+ 765 8.65 cadmium	49 In 114.82 1.7 157 3+ 2080 7.31 indium	50 Sn 118.69 1.8 232 4+ 2270 2+ 7.31 tin	51 Sb 121.75 1.9 631 3+ 1750 5+ 6.70 antimony	52 Te 127.60 2.1 450 2- 990 6.24 tellurium	53 I 126.90 2.5 114 1- 184 4.93 2 iodine	54 Xe 131.30 - -107 -108 3.52 xenon
55 Cs 132.91 0.7 28.4 1+ 678 1.90 cesium	56 Ba 137.33 0.9 725 2+ 1840 3.5 barium	57 La 138.91 1.1 921 3457 6.15 lanthanum	72 Hf 178.49 1.3 2227 4602 13.3 hafnium	73 Ta 180.95 1.5 2996 5425 16.7 tantalum	74 W 183.85 1.7 3410 5660 19.3 tungsten	75 Re 186.21 1.9 3180 5627 21.0 rhenium	76 Os 190.20 2.2 3045 5027 22.6 osmium	77 Ir 192.22 2.2 2410 4130 22.5 iridium	78 Pt 195.09 2.2 1772 4+ 3827 2+ 21.5 platinum	79 Au 196.97 2.4 1064 3+ 2807 1+ 19.3 gold	80 Hg 200.59 1.9 -39.0 2+ 357 1+ 13.6 mercury	81 Tl 204.37 1.8 304 1+ 1457 3+ 11.85 thallium	82 Pb 207.19 1.8 328 2+ 1740 4+ 11.4 lead	83 Bi 208.98 1.9 271 3+ 1560 5+ 9.8 bismuth	84 Po (209) 2.0 254 2+ 962 4+ (9.2) polonium	85 At (210) 2.2 302 1- 373 - 2 astatine	86 Rn (222) - -71 -62 4.4 radon
87 Fr (223) 0.7 (27) 1+ (677) 2.4 francium	88 Ra 226.03 0.9 700 2+ 1140 5.0 radium	89 Ac (227) 1.1 1050 3+ 3200 10.07 actinium	104 (261) Rf rutherfordium Ku kurchatovium	105 (262) Ha hahnium Ni nielsbohrium	106 (263)												

58 Ce 140.12 1.1 799 3+ 3426 6.67 cerium	59 Pr 140.91 1.1 931 3+ 3512 6.77 praseodymium	60 Nd 144.24 1.2 1021 3+ 3068 7.00 neodymium	61 *Pm* (145) - 1080 3+ 2460 7.22 promethium	62 Sm 150.35 1.2 1077 3+ 1791 7.52 samarium	63 Eu 151.96 - 822 3+ 1597 2+ 5.24 europium	64 Gd 157.25 1.1 1313 3+ 3266 7.90 gadolinium	65 Tb 158.93 1.2 1356 3+ 3123 8.23 terbium	66 Dy 162.50 - 1412 3+ 2562 8.55 dysprosium	67 Ho 164.93 1.2 1474 3+ 2695 8.80 holmium	68 Er 167.26 1.2 159 3+ 2863 9.07 erbium	69 Tm 168.93 1.2 1545 3+ 1947 9.32 thulium	70 Yb 173.04 1.1 819 3+ 1194 6.97 ytterbium	71 Lu 174.97 1.2 1663 3+ 3395 9.84 lutetium
90 Th 232.04 1.3 1750 4+ 4790 11.7 thorium	91 Pa 231.04 1.5 < 1600 5+ - 15.4 protactinium	92 U 238.03 1.7 1132 6+ 3818 4+ 19.1 uranium	93 *Np* 237.05 1.3 640 5+ 3902 19.5 neptunium	94 *Pu* (244) 1.3 641 4+ 3232 6+ - plutonium	95 *Am* (243) 1.3 994 3+ 2607 13.7 americium	96 *Cm* (247) - 1340 3+ - - curium	97 *Bk* (247) - - 3+ - 14 berkelium	98 *Cf* (251) - - 3+ - - californium	99 *Es* (254) - - - - einsteinium	100 *Fm* (257) - - - - fermium	101 *Md* (256) - - - - mendelevium	102 *No* (259) - - - - nobelium	103 *Lr* (257) - - - - lawrencium

EXERCISES TO BE SENT FOR CORRECTION

Exercise 1 A Brief History of Metals

1. Which three broad age groups did man go through?

2. Why was it an advantage to belong to a later age group?

3. Even though aluminum is the most plentiful metal in the earth's crust, man discovered other metals long before he discovered aluminum.

   (a) List at least three of these early metals.

   (b) Give one reason why they were discovered first.

4. What is an important feature of each of the three periods of the iron age?

Exercise 2 Iron

Give five reasons why iron is an extremely important metal.

Exercise 3 Prices and Production

1. Give one reason why entirely different methods have to be found to extract some metals from their ores.

2. Give five factors which influence the price of a metal.

3. Of the following metals: tin, nickel, copper, zinc, aluminum, and iron,

   (a) which is the most expensive?

   (b) which is the cheapest?

4. Canada is a leading producer of which three metals?

5. How many million tonnes of copper were produced in the world in 1974?

____

6. About how many metals are known today?

(a) 20
(b) 40
(c) 60
(d) 80
(e) 100 (_____)

Exercise 4 Properties

1. What is a simple definition of a metal?

____

____

____

2. All metals have

(a) very similar physical properties
(b) very similar chemical properties
(c) all of the above
(d) none of the above (_____)

Exercise 5 States

1. If you bend a metal bar, and it snaps right away, the metal would be in the

(a) crystal state
(b) plastic state
(c) liquid state (_____)

2. On a hot day, a chocolate bar is bent slightly but it doesn't break. The chocolate bar, however, is still cold enough so that all edges and any writing on it are clearly visible. The bar would be in a

(a) crystal state
(b) plastic state
(c) liquid state (_____)

3. If you put the bent chocolate bar in the fridge for 5 hours, what would happen if you tried to straighten it? Why?

Exercise 6

Obtaining Copper

1. What must be done to a large chunk of ore before the mineral can be floated off?

2. Balance the following equation for the roasting of Cu_2S.

$$\_\_\_\_ Cu_2S + \_\_\_\_ O_2 \rightarrow \_\_\_\_ Cu_2O + \_\_\_\_ SO_2 .$$

3. If you want to completely roast 318 g of Cu_2S, how many grams of oxygen do you need?

4. If you want to completely roast 477 kg of Cu_2S, how many <u>kilograms</u> of oxygen would you use?

5. (a) Balance the following equation for the reduction of Cu_2O to Cu.

$$\_\_\_\_Cu_2O + \_\_\_\_C \rightarrow \_\_\_\_Cu + \_\_\_\_CO$$

(b) From Lesson 8, what type of reaction is this?

____

6. How many grams of carbon would be required to reduce 143 g of Cu_2O?

7. (a) How many kilograms of carbon would be required to reduce 429 kg of Cu_2O?

(b) How many moles of carbon would this be?

8. Assume that one heaping shovelful of charcoal (carbon) weighs 10 kg. If you were the foreman, how many shovelfuls would you ask the worker to shovel into the container with 429 kg of Cu_2O? (In a practical operation, charcoal is cheap, but copper is expensive, so be a little on the safe side. Also do not expect the worker to know how much 0.537 shovelfuls is.)

END OF LESSON 8

LESSON RECORD FORM
1240 Chemistry 10
Revised 90/06

FOR STUDENT USE ONLY

Date Lesson Submitted

____

Time Spent on Lesson

____

(If label is missing
or incorrect)

File Number

____

Lesson Number __________

FOR SCHOOL USE ONLY

Assigned
Teacher: __________

Lesson Grading: ________

Additional Grading
E/R/P Code: ________

Mark: __________

Graded by: __________

Assignment Code: ______

Date Lesson Received:

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Lesson Recorded ______

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ALBERTA CORRESPONDENCE SCHOOL

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1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

(1) All pages are numbered and in order, and no paper clips or staples are used.
(2) All exercises are completed. If not, explain why.
(3) Your work has been re-read to ensure accuracy in spelling and lesson details.
(4) The Lesson Record Form is filled out and the correct lesson label is attached.
(5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

First Class

Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope. Correspondence lessons will travel faster if first-class postage is used.**

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

STEEL

A. Iron Ore to Cast Iron
    (1) Blast Furnace
    (2) What is Cast Iron?
B. Cast Iron to Steel
    (1) What is a Steel?
    (2) Bessemer Process
    (3) Open Hearth Furnace
    (4) Electric Furnaces
C. Iron and Steel
    (1) Iron and its Properties
    (2) The addition of Alloying Elements in Steel
        (a) Substitution Alloys
        (b) Interstitial Alloys
        (c) Pearlite Formation
        (d) Other Alloys
D. Different Types of Steels and their Properties
    (1) Carbon Steels
    (2) Tool Steels
    (3) Stainless Steels

A. Iron Ore to Cast Iron

How iron is found, and what is done to it, greatly affects its properties. Due to its great importance, it will be briefly discussed how pig iron is made from its ores.

(1) The Blast Furnace

The first step in the making of iron and steel occurs in the blast furnace. This can be up to 30 m high and 9 m in diameter.

Most iron ore today goes into the blast furnace to be made into pig iron. The most common iron ore is hematite (Fe_2O_3). Limestone and coke are added to the blast furnace along with the iron ore. The coke serves both as a fuel, and as a reducing agent, which helps to chemically reduce the iron ore to "pure" iron. The limestone combines with impurities of iron ore and forms a liquid slag which can easily be removed. However, limestone does not remove all impurities. Many more impurities will have to be removed later when pig iron is made into steel.

As with any operation, there are advantages and disadvantages to anything, and in this case, heating to the liquid state allows more carbon to be absorbed by the iron (from the coke). Although this pig iron has too many impurities to be used right away, the use of the blast furnace is still the best way to produce pig iron (the product of the blast furnace) in the long run. To give you an idea of the size of this operation, one can draw off up to 1100 tonnes of pig iron a day in a modern blast furnace.

Hot air is also added to the blast furnace. This hot air, with some chemical reactions, results in temperatures of 1930°C at the hottest part, or bottom of the blast furnace. Here is a pool of metal which is drawn off regularly.

On top of the liquid metal is a layer of liquid slag, since the slag is lighter than the metal. Slag is a combination of limestone, along with the impurities of the metal. This slag is drawn off through an outlet which is above that used for the molten iron. From the top, gases such as nitrogen and carbon monoxide are drawn off.

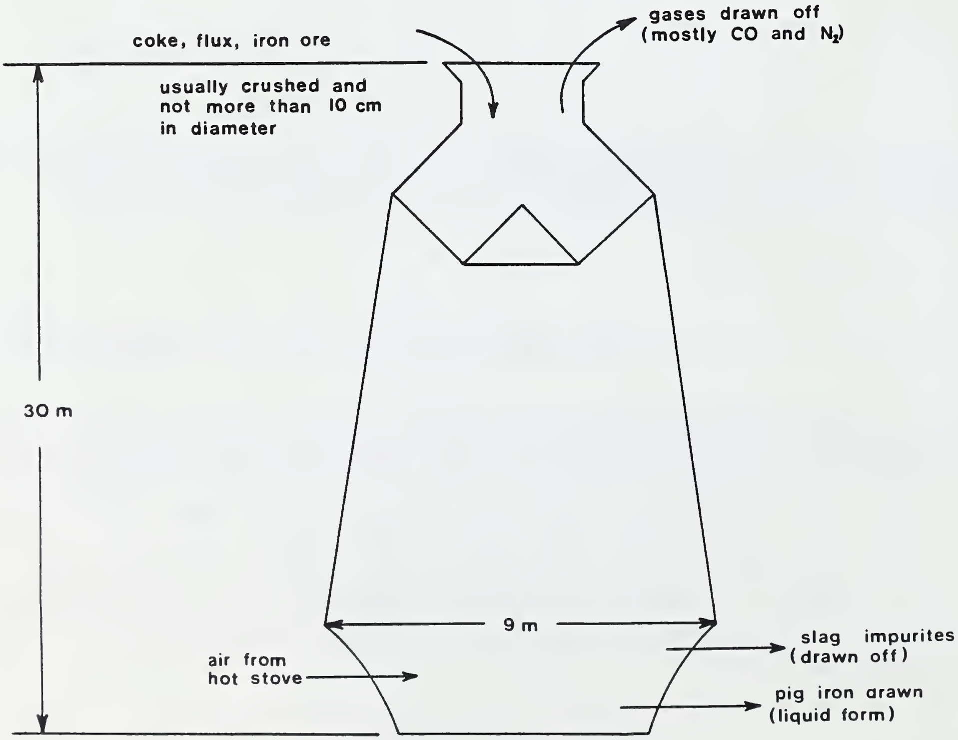


Figure 1: Simplified sketch of the blast furnace

(2) What is Cast Iron?

Cast iron, or pig iron, is the molten metal from the blast furnace. This iron will vary in composition, depending on the temperature, and on the composition of the original ore. The composition may be 90%-95% iron, 3.5% to 4.5% carbon, and smaller amounts of manganese, silicon, sulphur, and phosphorus. Of course, these elements will greatly affect the properties of iron, and much has to be done to purify the metal for any particular use. Not only will we remove many of the impurities, but other elements will be added to enhance certain properties. Incidentally, you may have noticed that many pans which are sold in stores are made of cast iron. It is the carbon which gives it a gray appearance.

Since much of the pig iron is used to make steel, steel making will be discussed next. There are a number of different kinds and different uses for steel, and it is natural that there would be a number of different processes used to make steel.

B. Cast Iron to Steel

(1) What is a Steel?

Plain carbon steel contains over 97% iron, and has small amounts of other elements. The most important additional element is carbon, which is usually less than 0.6% by weight, but a steel can never have more than 2% carbon, as contrasted with 4% carbon in pig iron. EVEN THOUGH THE CARBON CONTENT OF STEEL MAY BE LESS THAN 1%, THIS SMALL AMOUNT OF CARBON HAS A VERY SIGNIFICANT EFFECT ON THE PROPERTIES OF STEEL. Other elements, which were initially present in pig iron, may not be harmful in very small amounts, and they may even enhance certain properties of steel. However, most of the impurities have to be removed from pig iron.

Besides carbon steels, one may also have alloy steels, which would contain nickel, vanadium, chromium, tungsten and/or other elements in addition to iron. Properties of these steels will be covered later. Below, various methods of making steel from pig iron will be discussed.

(2) The Bessemer Process

The Age of Steel really began with the invention of the Bessemer Process in the middle of the 19th Century. The Bessemer converter is a large vessel which is lined with steel plates and heat resistant bricks. The usual size holds about 14 t of molten pig iron. The molten pig iron is poured through the top, and an air blast is forced through the molten mass. The oxygen combines with the impurities which results in the carbon being burned out, and which results in the removal of other impurities. (From the composition of cast iron, one can see that the limestone was not able to remove all impurities in the blast furnace). Extra materials may also be added to make the kind of steel desired. Today, however, this method of making steel is not used too frequently, since only high grade, low phosphorus ore can be used successfully with it.

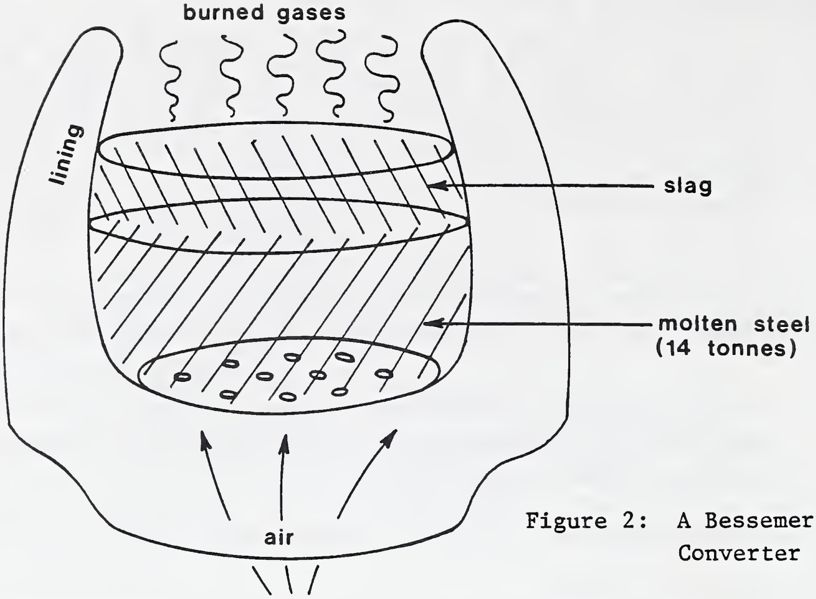


Figure 2: A Bessemer Converter

(3) Open Hearth Furance

Today, most of the steel is refined in open hearth furnaces which can refine ores of all grades. In this furnace, limestone, iron oxide, possibly scrap iron, and pig iron are added at various intervals. In this furnace, the hearth of the furnace, or the bottom, is exposed to flames, and preheated air is blown over the hearth to aid combustion. Temperatures that are reached may be 1650°C. Again, the limestone combines with the impurities to form a slag. The slag is the non-metallic product of the limestone and nonmetallic impurities in the metal. The lighter slag floats off from the top, and at the right time, molten metal flows into large ladles, which are receptacles for molten metal. At this time, appropriate alloying elements are added to make the kind of steel desired.

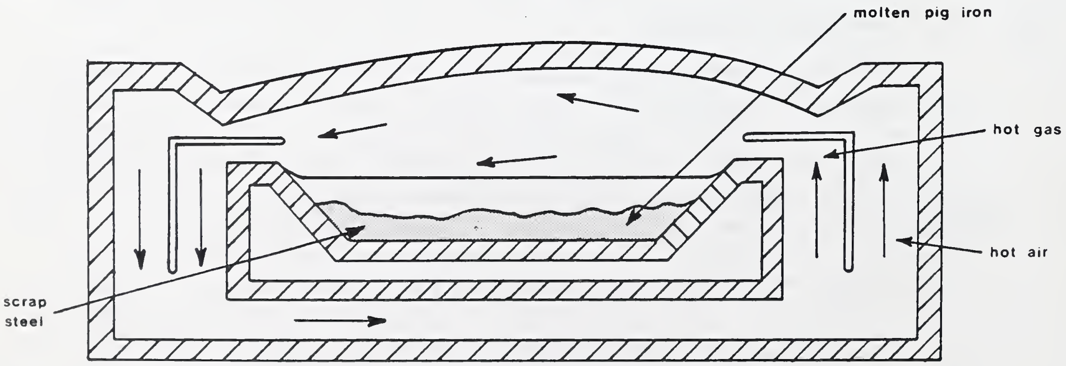


Figure 3: Open Hearth Furnace

(4) Electric Furnace

Although the open hearth furnace is used for most steels, steels of the highest quality are produced in arc furnaces. Many alloy steels such as stainless steel and tool steels need very high temperatures which can only be generated commercially with the electric furnace. In this furnace, electric arcs from carbon electrodes furnish the necessary heat. After the steel is molten, the appropriate alloys are added. With no flames, slags can be controlled better, and a steel of a higher quality can be made. The first question that may come to mind is why this method isn't used for all steels. The answer is that electric heat is more expensive than fuel heat, and, as with any process, the advantages must be weighed against the disadvantages.

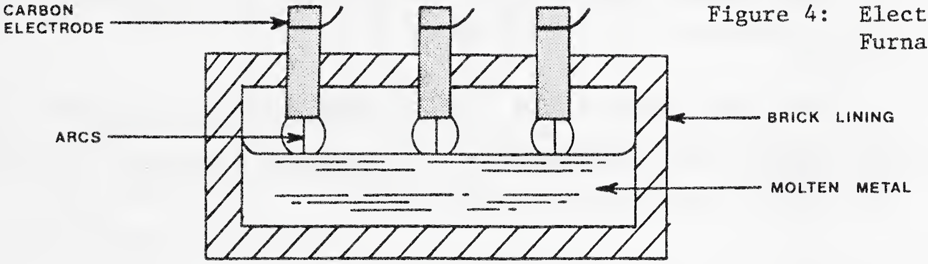


Figure 4: Electric Furnace

C. Iron and Steel

(1) Iron and its Properties

Although iron of over 99.9% purity has been produced in small quantities, it is not produced in this form commerically, partly because this is very difficult to do, and due to the fact that iron is more useful with some alloying elements. Iron melts at 1540°C, and it is ductile, strong, and magnetic.

Since alloying elements do not enhance the magnetic properties of iron, magnetism will be discussed at this point. Any piece of iron can be attracted by a magnet, and it can be magnetized. Iron can also be used in electromagnets, in which case the iron would act as a magnet when the current is on, and it would stop acting as a magnet when the current is off. All that an electro-magnet is, is an iron bolt, which has not been magnetized, but which has a coil of wire around it, and a power source. All electric motors and generators depend on the ability of iron to act as an electromagnet. Single magnets may be used in compasses, or to separate magnetic materials from nonmagnetic materials. Despite the good uses a magnet can be put to, the ability to be magnetized is not always a good quality as a watch may stop running if certain parts become magnetized. For this reason, many watches are made nonmagnetic. One way of doing this is to use stainless steel instead of iron. Incidently, iron also loses its magnetism at high temperatures.

Since iron is strong, cheap, abundant, and easy to reduce from its ore, it is the most important metal. However, iron almost always has carbon and other alloying elements in it, which would make it a steel. In the next section, we will see how the various alloying elements combine with iron to produce steel.

(2) The Addition of Alloying Elements in Steel

Steel is an alloy of iron and carbon and other elements. Even though the carbon content is extremely low, it has a marked affect on the properties of steel. Commercial steels always contain manganese, silicon, phosphorus, and sulfur and possibly residual amounts of other elements. The effects of these other alloying elements is generally small. The effects of the carbon however, is not very straightforward, due to the many ways in which we can combine carbon with iron. Heat treatments have a great effect on how the carbon combines with the iron.

Before we discuss how other elements alloy with iron, we will have to know how the atoms in "pure" iron are arranged. Although the diagram is not strictly accurate, it will adequately illustrate how other atoms can alloy with iron.

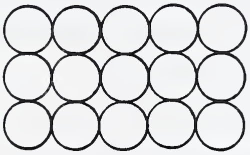
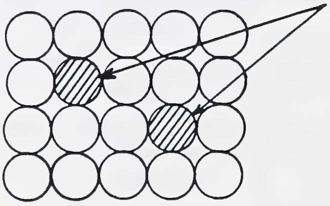


Figure 5: "Pure" Iron

(a) Substitution Alloys

Technically speaking, an alloy is defined as two or more metallic elements, or metallic and nonmetallic elements which are soluble in each other when molten, and which do not separate into distinct layers when solid. It is easy to see that an atom which is much heavier than the iron atom or one that is much larger than the iron atom will not comfortably fit in with the other iron atoms, and may sink to the bottom if iron is in the molten state, however, in the crystal structure, many atoms can readily substitute for an iron atom without doing too much damage. The diagram below illustrates how one atom substitutes itself for an iron atom.



Chromium atoms substitute for iron atoms (only if no carbon is present)

Figure 6: Substitution Alloy

(b) Interstitial Alloys

Small atoms, such as the carbon atom, can enter the solid interstitially. The carbon atoms are much smaller than the iron atoms, therefore they cannot replace an iron atom. What they do, is go in between iron atoms. Since the space between iron atoms is a bit too small for the carbon atom to fit in comfortably, very few carbon atoms do this. At high temperatures, up to two per cent by weight fit in interstitially, but at room temperature a very small amount of the carbon can fit in, in this manner.

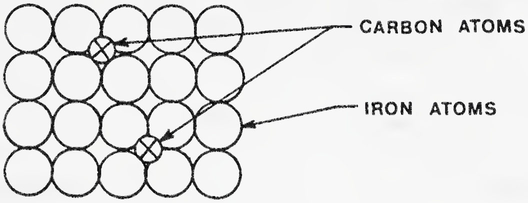


Figure 7: Interstitial Alloy

As will be seen in the next section, most carbon atoms enter the solid iron in a rather unique way.

(c) Pearlite Formation

Since very few carbon atoms can enter the iron interstitially, a new complex, cementite, is formed. Cementite is Fe_3C, which means that in a small grain of cementite, one has three iron atoms for every carbon atom.

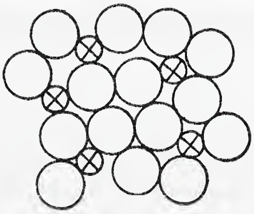


Figure 8: Cementite

The actual structure of cementite is much more complex than that which is illustrated, but the general picture should be clear. Now, in steel, one never has that much carbon to form cementite throughout the structure. However, there is too much carbon for it all to enter interstitially. What happens then, when you have one per cent carbon, is that you have sections of the piece of steel which are composed of cementite, and you have other sections which have an extremely small amount of carbon. This portion, that can dissolve a small amount of carbon at room temperature is called ferrite. Steel, containing cementite mixed in with ferrite, is called pearlite.

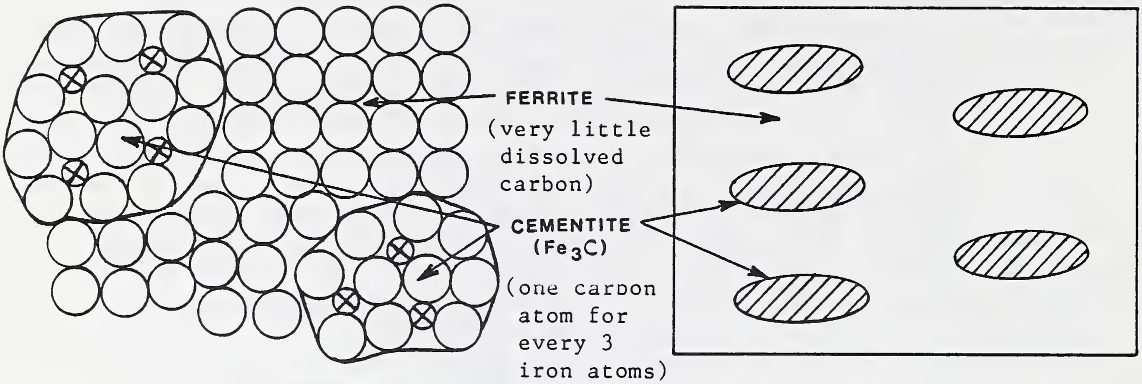


Figure 9: Pearlite

A(atomic scale)

B(under a microscope) 3000 time magnification

In the above diagram, note the very odd carbon atom in ferrite, however, in cementite, the ratio of carbon atoms to iron atoms should be one to three. In an actual case, one could never see the individual atoms, however the diagram on the right illustrates what one might see on a flat piece of steel that has been magnified about 3000 times.

Many other elements which are added to steel may form carbides which dissolve in the iron just like cementite does. For certain steels, instead of having iron carbide or cementite (Fe_3C), you may have a chromium carbide ($Cr_{23}C_6$) or a vanadium carbide (V_4C_3) or other carbides.

2 (d) Other Alloys

Now, cementite and ferrite and pearlite have been discussed. Keep in mind that pearlite is no new structure, but it is just a combination of cementite and ferrite. There are also other structures, some of which will be discussed further in the next lesson. Depending on how high one heats something, and on how fast one cools it, one may get austenite, pearlite, martensite or tempered martensite. The four things mentioned in the previous sentence are all forms of steel, and all can have the same composition. For example, one could have an alloy of 0.8% carbon by mass, and 99.2% iron by mass, yet the alloy can be either pearlite, austenite, martensite or tempered martensite. The name given to the steel depends on the exact way in which the carbon and iron atoms are arranged.

In whichever way the carbon atoms might be arranged in steel, one never has two or more carbon atoms side by side. In cast iron, however, one often has clusters of carbon atoms, and this has a weakening effect on the structure. As a result, cast iron is always weaker than steel. As an analogy, think of the seeds in a watermelon, and how they weaken the structure.

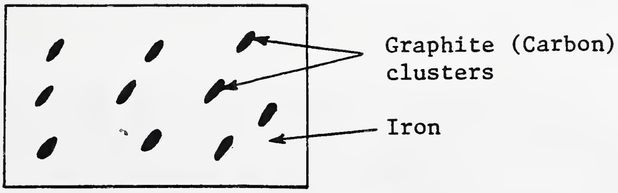


Figure 10: Cast Iron

D. Different Types of Steels and their Properties

There are different types of steels depending on which alloying elements are added to the iron and carbon. Steels which just contain carbon and iron are called carbon steels. Alloying elements which are generally added to tool steels usually affect the physical properties such as making the steels harder and stronger. Other elements, such as chromium, are added to stainless steels to change the chemical properties of the steel in order to make it more corrosion resistant. Although commercial steels almost always contain manganese, silicon, sulfur, and phosphorus, these elements will not be of concern in the following section, since these elements are present in too small amounts to have much effect on the properties of the steel.

(1) Carbon Steel

Carbon steel is a steel in which carbon is the chief alloying element. Carbon steel is arbitrarily divided into low carbon, medium carbon, and high carbon steel. Each type of steel has its own particular advantages and disadvantages, along with its own particular uses.

Low carbon steel has from 0.10 to 0.25% carbon. It is commonly used in the manufacture of chains, nails, pipes, and structural shapes.

Medium carbon steels have from 0.25 to 0.55% carbon. Giving the carbon content alone can be a bit misleading, since two steels with the same carbon content can have greatly different properties, depending on the heat treatment they were given. However, medium carbon steel is used for axles and crank shafts.

High carbon steels have from 0.55 to 1.00% carbon. The higher carbon content results in the steel having a decreased machinability, poor formability, and poor weldability. However, these steels can be made harder than any other carbon steel. However, the hardness is only on the surface since it is very difficult to make the whole piece of steel hard throughout, with just carbon, unless you have a very thin piece of steel. High carbon steels are often used for chisels, knives, and files. If one wants a larger piece of steel to be harder throughout, various alloys must be used.

(2) Tool Steels

Although high carbon steels are often used for tools, and steels discussed here are steels with other alloying elements in addition to carbon. The carbon content in tool steels is usually 0.60% anyway, since this amount of carbon enables one to make the steel very hard. Other alloying elements also make the steel harder since they combine with carbon in a similar way that iron does when cementite or iron carbide (Fe_3C) is formed. The vanadium carbide and the chromium carbide is much harder than the iron carbide however.

Besides making the steels harder, alloying elements also help to change other physical properties such as the melting point. For example, the melting point of tungsten is much higher than that of iron. What do you suppose the melting point would be if some tungsten were alloyed with iron? The melting point itself is not too important however, since although pure iron melts at 1540°C it loses its usefulness as a tool well below this temperature. Other elements also help the iron to remain sufficiently strong at much higher temperatures. One useful alloy which has been used as a high speed cutting tool, and which does not lose its hardness even when it is red hot, is a steel which has been alloyed with tungsten or molybdenum or both, and it may also contain chromium, vanadium, or manganese.

Most of these elements combine with carbon to form a carbide, instead of just existing as individual atoms in iron. The ratio of these atoms with carbon atoms varies, but it would be different from the iron carbide (Fe_3C) or cementite, which we discussed earlier. Chromium increases the hardenability, or the ability to make steels harder, and its carbide has excellent wear resistance. Vanadium forms the hardest carbide of all, and has the greatest wear resistance, however due to its expense, it isn't used unless such a high hardness is necessary.

Of course, there are many different combinations and amounts in which the various elements can be added to iron to give it the desirable properties. These alloying elements can have an effect on the properties of steel if they are present in amounts of less than 1%, however, the amount of any alloy may be 12% or higher, depending on the particular use required.

Steels which retain their strength in the temperature range of 760 - 980°C have also been developed. Alloying elements which help steel attain a high strength in this temperature range are nickel and cobalt. However, when one gets above this range, the best steels are not suitable, and entirely different metals will be needed, as will be seen later.

(3) Stainless Steels

You are probably familiar with rust, which is an iron oxide layer on iron. In order to combat this, and to change the chemical properties of iron, other elements are often added. The most important element which is added to combat corrosion is chromium, although nickel and manganese may also be present. Just as there are many combinations in which various elements can be added for various tool steels, it is also the case that various combinations of elements can be added for various types of stainless steels. Although there are exceptions, many stainless steels have not more than 0.20% carbon, 1 - 2% manganese, 12 - 20% chromium, and 0 - 14% nickel. You may wonder why so many substances are often added. Although chromium is the most important element for corrosion resistance, a chromium steel does NOT do well against chloride ions. Nickel and molybdenum improve resistance against chlorides, and in addition, molybdenum does particularly well against sulfuric and sulfurous acids.

Besides being corrosion resistant, stainless steels are also required to have reasonable mechanical properties. As was stated before, chromium also increases the hardenability of steels, and other elements may also be added for certain mechanical properties.

EXERCISES TO BE SENT FOR CORRECTION

Exercise 1 Producing Cast Iron

1. Which four things are put into the blast furnace?

____

____

____

____

2. Which three things leave the blast furnace?

____

____

____

3. What is the purpose of the limestone?

____

____

____

4. One reaction that takes place in the blast furnace is

$$\_\_\_\_ Fe_2O_3 + \_\_\_\_ C \quad ' \quad \_\_\_\_ Fe + \_\_\_\_ CO$$

Balance this equation.

5. How many kilograms of carbon are required for 160 kg of Fe_2O_3. (The ratios are the same, whether you work in grams, kilograms, or tonnes.)

6. A modern blast furnace can produce 1.10×10^3 t of pig iron (Fe) in a day. This is equal to 1.1×10^6 kg or 1.1×10^9 g of Fe.

(a) How many kilograms of charcoal (carbon) would a day's operation require?

Hint: First calculate the number of moles of Fe produced.

$$\frac{1.1 \times 10^9 \text{ g}}{56 \text{ g/mol}} = 1.96 \times 10^7 \text{ mol.}$$

Next, using your <u>balanced</u> equation on page 12, calculate by means of ratio and proportion the number of moles of charcoal that are required. (If necessary refer to examples on page 2 of Lesson 8.) Finally, convert moles to grams and then to kilograms. Follow the steps given below, please.

$$\frac{1.96 \times 10^7 \text{ mol Fe}}{2 \text{ mol Fe}} : \frac{x \text{ mol C}}{3 \text{ mol C}}$$

x = ______________________ mol of carbon.

Finally convert moles to grams and then to kilograms.

______________________ g of carbon

______________________ kg of carbon

You may also solve the above problem using mass ratios as follows:

$$\frac{1.1 \times 10^6 \text{ kg Fe}}{2(56)\text{g Fe}} : \frac{x \text{ kg C}}{3(12)\text{g C}}$$

x = ______________________ kg.

(b) How many <u>tonnes</u> of charcoal would a day's operation require?

$$(1 \text{ t} = \frac{1.1 \times 10^6}{1100} = 1 \times 10^3 \text{ kg})$$

7. On a given week, a blast furnace had the following output of pig iron. On day one, the output was 1.1×10^6 kg. On day two, the output was 1.43×10^6 kg since a high quality of ore was used this day. On day three, the output was $1.095\ 437\ 9 \times 10^6$ kg. On day four, the machine was down for repairs part of the time, and only 5.396×10^5 kg was produced. On day five, things still didn't go properly, and only 7×10^4 kg were produced. Using the proper number of significant digits, calculate the <u>total</u> output for these five days.

8. (a) Using either method shown in Exercise 6, calculate the mass of charcoal that is required to produce 5.396×10^5 kg of iron in day four. Refer to your balanced equation on page 12 to obtain proper ratios, please.

   (b) On day five, 7×10^4 kg of pig iron (Fe) were produced. Calculate the mass of iron ore (Fe_2O_3) that was required.

9. (a) From question (3), you know what limestone ($CaCO_3$) does. One reaction that takes place is

$$CaCO_3 + SiO_2 \longrightarrow CaSiO_3 + CO_2$$

(the reason this reaction has to take place is that SiO_2 has a melting point higher than that of iron, and a liquid slag is desired.)

Is the equation balanced? ______________________

If it is not, balance it.

(b) From Lesson 8, what type of reaction is this?

____

10. Suppose that your particular ore has 25% of SiO_2 in it, so for every 1000 kg of ore you have 250 kg of SiO_2. Assuming that all of your limestone, $CaCO_3$, combines with the SiO_2 in the ore and with no other impurities, calculate the mass of $CaCO_3$ that is required to combine with 4.8×10^4 kg of ore.

(**Hint:** 25% of 4.8×10^4 kg = 1.2×10^4 kg of SiO_2 that are actually reacting, agree?)

If you worked through Exercise 6 on page 12, you should have no difficulty solving the above problem. Your equation is given on top of this page.

11. (a) In the previous two equations, CO and CO_2 were given off. Which of these gases occurs naturally in the environment?

____

(b) Which one is a poison?

____

(c) If the gas which occurs naturally in the environment is produced in such large quantities that the plants are unable to use up the excess, do you think that this gas should also be considered a pollutant?

____

Exercise 2 Producing Steel

1. Which element is important in determining whether or not something is a cast iron or a steel?

____

____

2. Iron, which has 1% carbon, can have greatly different properties than iron with no carbon in it. True or False?

3. Besides carbon, name four other elements which may be added to form a certain type of steel.

____

____

____

____

4. (a) What is the purpose of the Bessemer converter?

____

____

(b) What is its capacity?

____

____

5. What is used to refine most of our steel today?

6. Which types of steels are produced in the electric furnace?

7. Give two reasons why the electric furnace is used to make these types of steels, instead of the open hearth furnace.

8. The following reactions occur in the open-hearth furnace. In each case write the balanced equation, and state what type of reaction you have.

(a) iron II sulfide + iron II oxide → iron + sulfur dioxide

balanced equation: ____________________

type of reaction: ____________________

(b) calcium oxide + diphosphorus pentoxide → calcium phosphate

balanced equation: ____________________

type of reaction: ____________________

Exercise 3 Iron

To avoid watches from becoming magnetized, manufacturers may use ______________ instead of iron.

Exercise 4 How Elements Alloy with Iron

1. Which element must be added to iron to get a steel?

____

2. Does a fixed percentage of this element (in #(1)) result in the same properties of steel at all times? Why or why not?

____

____

____

3. If an atom has the same size and weight as an iron atom, it can enter the iron interstitially. True or False?

____

4. At room temperature, why does an almost negligible number of carbon atoms enter the iron interstitially?

____

____

5. Which two distinct phases are in pearlite?

____

____

6. Which of the phases in #5 has a much higher carbon content?

____

7. If 0.50% carbon makes a steel much stronger, why isn't cast iron still much stronger than steel, since it has 4% carbon?

____

____

____

8. If you have an atom the same size as iron, how will it enter a sample of pure iron?

   (a) interstitially
   (b) substitutionally
   (c) as a carbide (____)

Exercise 5 Carbon Steel

1. If you had 1000 g of steel which consisted of 994 g of iron, and 6 g of carbon, this would be considered a

   (a) low carbon steel.
   (b) medium carbon steel.
   (c) high carbon steel. (____)

2. The hardest steel is

   (a) low carbon steel.
   (b) medium carbon steel.
   (c) high carbon steel. (____)

3. Since the carbon atoms prevent the iron atoms from sliding over each other in cementite, and since the carbon atoms have next to no effect on ferrite, which would be harder: cementite or ferrite?

   ____

4. Which is more ductile: cementite or ferrite?

   ____

5. Austenite, pearlite, martensite and tempered martensite are all different forms of steel, and it is possible that all can have the identical composition. True or False? (____)

Exercise 6 Tool Steels

1. What is the general difference between a tool steel and a low carbon steel?

   ____

   ____

2. List *five* of the common alloying elements used to make tool steels.

   ____

   ____

3. What is the one most important property of tool steels?

   ____

4. Which of the following metals is NOT used to make steels harder: tungsten, molybdenum, copper, vandium, nickel, cobalt.

____

Exercise 7 Stainless Steel

1. Which is the most important metal for corrosion resistance?

____

2. How much chromium does a stainless steel commonly have:

(a) 2%
(b) 18%
(c) 30%
(d) 50% (_____)

NOTE: IF YOU HAVE BORROWED A BEAM BALANCE, A SET OF WEIGHTS, A THERMOMETER AND A GRADUATED CYLINDER, PLEASE RETURN THESE TO THE ALBERTA CORRESPONDENCE SCHOOL IF YOU HAVE NOT ALREADY DONE SO.

END OF LESSON 10

LESSON RECORD FORM

1240 Chemistry 10

Revised 90/06

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(If label is missing
or incorrect)

File Number ____________________

Lesson Number __________

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Assigned Teacher: __________

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ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

(1) All pages are numbered and in order, and no paper clips or staples are used.
(2) All exercises are completed. If not, explain why.
(3) Your work has been re-read to ensure accuracy in spelling and lesson details.
(4) The Lesson Record Form is filled out and the correct lesson label is attached.
(5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

First Class

Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope. Correspondence lessons will travel faster if first-class postage is used.**

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

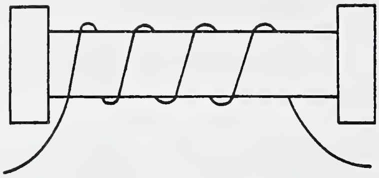
OTHER METALS

A. General Discussion

Despite the fact that more iron and steel is produced than all other metals combined, there are many other metals which have properties which steel does not have. Even though there are about eighty metals on the periodic table, we still need alloys for many different applications.

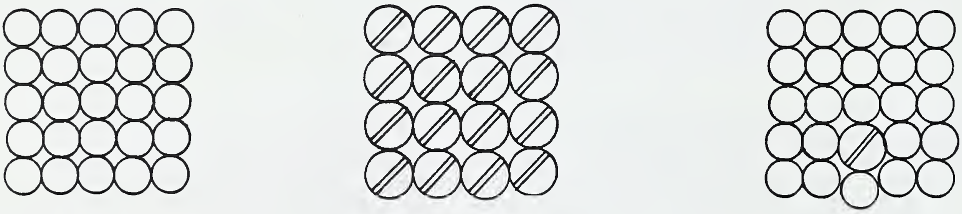
We will define an alloy as two or more metals which are soluble in each other when molten, and which do not separate into distinct layers when solid. Pairs of metals which DO mix in all proportions are copper-nickel, silver-gold, nickel-chromium, and many others. Other metals mix in each other only when you have certain (usually low) amounts of one metal present. Still other metals do not mix at all. If you melt aluminum with lead for instance, the lighter aluminum will float on top, and the lead will be at the bottom, just like oil and water, and no alloy would be formed. An important feature about alloys is that their properties are often not in between the properties of the two combining metals, but their properties may be better than that of either contributing metal alone. For example, bronze is much harder than either copper or tin alone would be.

B. Copper



Copper is the most important metal outside iron and its compounds. The reason for its great importance is that copper is an excellent conductor of electricity. Copper is what makes generators, motors, and many other electrical instruments workable. In lesson nine, a metal was defined to be an element which readily loses electrons. Since copper loses its electrons much more readily than most other metals however, it is used extensively for wiring.

In addition to its use as a conductor, copper is also used as copper tubing, either by itself, or alloyed with other elements. As an illustration of how alloys greatly improve the properties of a metal, consider how tin affects copper. Copper is harder than tin, but when you add only 5% tin to copper, you get a bronze which is twice as hard as copper.



A - Pure Copper B - Pure Tin C - Bronze

From the above diagrams, it can readily be seen that if all atoms are the same size it can take a certain amount of force to make the atoms slide over each other, however, with bronze, the odd tin atom is a great hindrance when the copper atoms want to slide over each other. Bronze, as well as brass, are substitutional alloys in which the tin or zinc atom substitutes for a copper atom.

Brass is a mixture of copper and zinc, however these two metals do not mix well in all proportions. As it turns out, due to differences between zinc and copper atoms, only about 40% of the zinc atoms can conveniently substitute for copper atoms. Increasing the amount of zinc has a favorable effect on some properties, and an unfavorable effect on other properties. For example, the tensile strength, yield strength, and hardness are increased with increasing amounts of zinc, whereas electrical conductivity and thermal conductivity are decreased with increasing amounts of zinc. Another example as to how impurities affect conductivity is that copper, which has only 0.3% arsenic, has its conductivity decreased by a half.

(1) Conductivity

Although all metals are relatively good conductors, conductivity will be discussed at this point since copper, aluminum, and silver are some of the better conductors of electricity. Incidentally, silver is a better conductor of electricity than copper, but the reason that copper is used is that copper costs about $1.43/kg, whereas silver is worth $140.00/kg.

Electrical conduction is basically a movement of outer electrons, and the easier that this occurs, the better the metal conducts.

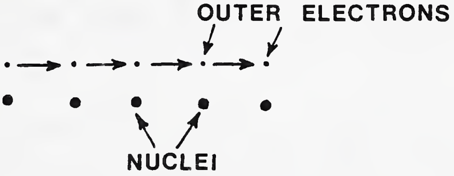
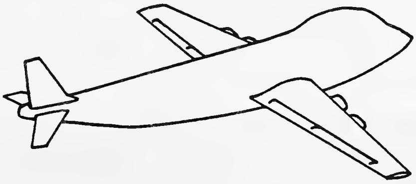


Figure 2: Conduction

In conducting electricity, the outer electron goes from one atom to another. Electrical conduction is extremely fast, however one electron does not move very fast. What makes conduction fast is that the first outer electron affects the second outer electron, which in turn affects the next outer electron, etc. Perhaps an analogy can be drawn with a set of bowling balls in a line. If ten bowling balls were in a line and touched each other, and a slowly moving additional bowling ball rolled to touch the first bowling ball, the last bowling ball would take off almost immediately. Even though the last bowling ball would not be travelling any faster than the first one was, it would get to the end of the line sooner than the first moving ball would have if it had not come in contact with any bowling ball.

Of course one must keep in mind that a copper wire has billions of atoms, and conduction would be much more complicated than was indicated.

C. Aluminum



Aluminum is the most plentiful metal in the earth's crust. Some properties of aluminum are that it is malleable and ductile, and it is a good conductor of heat and electricity. It can be readily cast, rolled, forged, extruded, or drawn. Other qualities of aluminum are that it is extremely light, about one third as heavy as steel, and it is very corrosion resistant as it forms a thin oxide film very quickly, which does a great deal to prevent further oxide from forming.

Naturally, a metal with all of these good qualities should be put to many good uses. Being a good conductor of electricity makes it possible to use aluminum to transmit electricity over long distances. Due to its lightness, aluminum is often used as parts for airplanes where a lighter metal would result in a great saving in fuel costs. Other common uses of aluminum are in cooking utensils, aluminum foil, and window frames.

Although extractive metallurgy is not an important part of this course, it is interesting to note that aluminum was not used for much in earlier days since it was extremely difficult to obtain aluminum from its ore, and therefore aluminum was very expensive. When a cheaper way was found to extract aluminum from its ore in 1886, the price dropped from $17.50/kg to $0.33/kg. Another metal which had a similar problem was titanium.

Although aluminum has many good qualities, it has a drawback in that pure aluminum is relatively weak. This is not so serious however, since aluminum can readily be alloyed with other substances which would greatly increase its strength. As one example, adding 4% copper can greatly increase the hardness and strength of aluminum, and at the same time, this alloy would not be much heavier than pure aluminum.

D. Noble Metals

Noble metals include gold, silver, platinum, iridium, osmium, palladium, rhodium, and ruthenium. These metals are particularly noted for their corrosion resistance, and they do not need an oxide film, as is the case with aluminum and other metals, for their corrosion resistance.

(1) Silver

Silver is the best conductor of heat and electricity, and it is the second most malleable and ductile metal. It is so malleable that if all pages of Encyclopedia Britannica were printed on the thinnest silver leaf, the pages would be only five-eighths of a centimetre. Silver plating of tableware is a very important use of silver. Silver is also used extensively in the photography industry. Other uses include bearings, brazing alloys, electric contacts, jewelry and coinage.

Silver is also alloyed with other metals such as copper to make it tougher and harder. Sterling silver, for example, has 92.5% silver and 7.5% copper.

(2) Gold

Gold is not attacked by air, water, or any of the common acids. Gold is used in jewelry, and once was frequently used in coinage. Gold plating is also so thin that 300 000 sheets would be only 2.5 cm thick. Pure gold is yellow, but white gold is an alloy of gold with usually another noble metal. Due to its softness, gold is often alloyed with copper to make it harder.

(3) Platinum

Platinum is a heavy metal which is very resistant to corrosion. It is also so ductile that it can be drawn into a wire so fine that it cannot be seen with the unaided eye. Although most metals are good conductors, platinum has a relatively high electrical resistance. Therefore it is used in making electric furnaces, resistance thermometers, spark plug tips, and X-ray and radio tubes. In addition to also being used for jewelry, platinum is used in the chemical industry as a catalyst in making sulfuric acid and nitric acid.

E. Refractory Metals

Refractory metals are metals with high melting points which are above 1870°C, and include tungsten, vanadium, chromium, columbium, molybdenum, tantalum, and zirconium. Tantalum is used with various electrical and chemical equipment, and it used to be put into incandescent lamps before it was replaced by tungsten. Zirconium is frequently used in nuclear technology. An interesting fact about molybdenum is that it is stiffer than steel, and it would therefore deflect much less under a given load. Molybdenum and vanadium are often alloyed with steels to make high speed tool steel alloys. The refractory metals, along with titanium, may be used more in the future when nuclear rocket engines and other things are built which work at temperatures of above 1000°C.

Chromium, as you may recall from lesson ten, is very highly corrosion resistant. As a result, it is often plated with other metals. Chromium is the most important alloying element of stainless steels, which often contains 18% chromium. Chromium also makes steels harder.

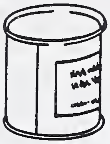
(1) Tungsten

Tungsten is the most important refractory metal and has the highest melting point, 3410°C. Compare this to iron whose melting point is 1540°C. Due to its high melting point, tungsten is often used in electric light bulbs. As mentioned before, much tungsten is alloyed with steel, which helps steel remain hard, even when it is red hot. As a result, boring and cutting machines do not lose their keen edges as they become hot.

(2) Powder Metallurgy

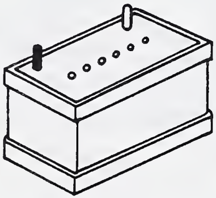
Many metal products are made by casting and possibly also machining the metal into the desired shape. However it is extremely difficult to cast a metal like tungsten, due to its high melting point. Tungsten, along with other refractory metals, are often produced in certain shapes by using the powder of the metal, and then by heating the compressed metal powder in a sintering furnace. In this sintering furnace, the metal powders are joined together. With tungsten, the powder is produced by reducing WO_3 powder to the metal powder. Nickel powder is often produced by pouring the liquid metal into water.

F. Tin



Tin is extremely resistant to rusting, and it is also resistant to many acids. As a result, it is often plated on sheet iron to protect it from corrosion. "Tin" cans are mostly iron, which have a thin layer of perhaps 38/1 000 000 cm of electrolytically plated tin on them. Another important use of tin is in soldering, which is a method of joining metals.

G. Lead



Lead is a soft and malleable metal. Some of its uses include storage batteries, cable covering, and it is used in a compound for leaded gasolines. Since it has the low melting point of 327°C, lead can be used in soldering. In early days, lead was used by the Romans for water pipes.

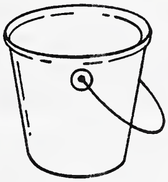
The overall reaction in the lead storage battery is ($Pb + 2HSO_4 + 2H + PbO_2 \rightarrow 2PbSO_4 + 2H_2O$.) The negative and positive charges indicate the presence of ions in the solution. We will not go into detail here, but note that this reaction is not covered by the types of reactions you studied in Lesson 8. It turns out that the above reaction is an oxidation-reduction reaction.

In the wrong place, lead can be a health hazard. For instance tetraethyl lead ($Pb(C_2H_4)_4$) is often added to gasoline to raise the octane rating of the gasoline. This lead soon gets into the air, and your health can be impaired if you breath air with a high concentration of lead over a long period of time. In the body, lead can cause anemia, headaches, brain damage, or even death if the dose is large enough.

Air pollution is often measured in parts per million, and in the case of gases, a 1 ppm concentration means 1 molecule of the contaminant to a total of a million molecules. The danger with lead poisoning is not just the amount you get, but how you get it. The lead in the air that you breathe has a much worse effect on you than the lead that may be in your food or water.

You may think that some concentrations are extremely small, but the danger is that you may breath bad air in, but don't breathe the bad lead out nearly as fast, so the concentration of lead in your lungs will soon be much higher than the outside air.

H. Zinc



Zinc forms a protective coating, making it somewhat corrosion resistant, however, if you have zinc and another metal side by side, the zinc will often be the first to corrode away. When steel is coated with zinc, as with a galvanized pail, any corrosion that would occur, would attack the zinc first, thus protecting the steel. Zinc is also used in the cathode of dry cells. Brass, which is an alloy of zinc and copper, has many uses, including water pipes.

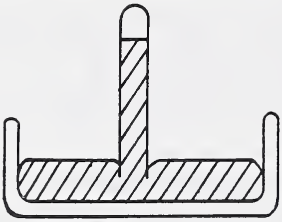
I. Nickel

Nickel is very corrosion resistant, and for this reason, it is used extensively for plating. Nickel is also a very important alloying element. Monel metal, which is 70% nickel, 28% copper, and 2% iron is a strong acid resisting alloy. Many nickel alloys such as alnico are more magnetic than iron. Many alloys, which are very strong at high temperatures, are also produced, in which nickel is often used as a major alloying element, and precipitation hardening is a major contributor to greater strength. Stainless steels often have 8% nickel in them.

J. Titanium

Titanium ores are very abundant, and the metal has a high strength, considering how light it is. It also has a high melting point and excellent corrosion resistance. Corresponding with aluminum, its corrosion resistance is due to the titanium oxide film on the surface. This metal would be used much more if it were not for a serious drawback, — it is very difficult to obtain titanium from its ore. Also, one must have inert environments before one can melt or weld it, since it would absorb gases at high temperatures and become brittle. An important use of titanium is in aircraft and missiles.

K. Mercury



Mercury is the only metal which is a liquid at room temperature, since its melting point is -40°C. It is so heavy that a chunk of iron would float in a pool of mercury. Mercury (quicksilver) is commonly used in thermometers, barometers, and other scientific instruments. The reason mercury works in thermometers is that it possesses the same property as iron and most other metals in that it expands on heating and contracts on cooling. Antimony is an exception since it expands when it solidifies. One should be careful when working with mercury or its compounds, as they may be poisonous.

Mercury has been in the news lately as an environmental hazard. Mercury can exist in several forms, and some forms are more dangerous than others. The alkyl mercury form such as CH_3Hg^+ and $(CH_3)_2Hg$ are particularly dangerous and can cause brain damage. Since these forms of mercury are not readily broken down by the body they remain in the body for a long time.

It is important to realize that you do not have to eat large amounts of contaminated fish in order to endanger your health. Fish will be removed from the shelves if the concentration of mercury is over 1 ppm. One ppm (part per million) as far as solids and liquids are concerned is a weight relationship. Thus tuna fish, with a mercury concentration of 1 ppm means that in a million grams of the fish, you have 1 g of mercury.

Since the problem results because the mercury is concentrated in your body and does not get eliminated nearly fast enough, both factors play an important role, — the concentration of the mercury in the fish, — and how much fish you eat every week.

You may have heard of the case in Minamata, Japan, where many fisherman and their families suffered mercury poisoning. Many died, and many others suffered all kinds of permanent disabilities.

Sample problem: Fishing restrictions will be imposed if the fish are found to have a high mercury concentration. Supposing that you have a 10 kg fish with a mercury concentration of 2 ppm.

(a) What is the mass of the mercury in the fish?

(b) How many moles of mercury are there?

(c) How many atoms of mercury do you have?

Solution

(a) 2 ppm means 2 g in 1 000 000 g
10 kg is 10 000 g
Using ratio and proportion, we get

$$\frac{2\ \text{g}}{\text{X g}} = \frac{1\ 000\ 000}{10\ 000}$$

Cross multiplying we get:

$$1\ 000\ 000\text{X} = 20\ 000$$
$$\text{X} = \frac{20\ 000}{1\ 000\ 000} = .02$$

So the 10 kg fish has 0.02 g of mercury.

(b) 1 mol of mercury weighs 200 g

$$.02\ \text{g} = \frac{.02}{200} = \frac{2 \times 10^{-2}}{2 \times 10^{2}} = 1 \times 10^{-4}\ \text{mol}$$

(c) 1 mol of mercury has 6.02×10^{23} atoms

10^{-4} mol of mercury has 6.02×10^{19} atoms

L. Arsenic

The most important feature of this gray, brittle metal is that it is poisonous. Due to its poisonous nature, it is used against insect pests.

M. Alkali Metals

The alkali metals, which are found on the left side of the periodic table, include lithium, sodium, potassium, rubidium, cesium, and francium. From a chemical point of view, francium is the most metallic metal, since it loses its outer electron easier than all other metals. (Recall the definition of a metal from lesson nine.) However, the alkali metals are not important structural materials as they are all soft and have low melting points. As a matter of fact, cesium and francium melt just a little above room temperature. The major importance of alkali metals lies in the compounds they form with other elements. Two very important compounds of sodium are sodium chloride, which is ordinary salt, and sodium hydroxide, which is an important base. Potassium salts are very important for plant growth.

N. Alkaline-earth Metals

The alkaline-earth metals are in the second group in the periodic table, and include beryllium, magnesium, calcium, strontium, barium, and radium. Along with alkali metals, they are generally white in color, and are not important structural materials, except for magnesium, although their salts are important. Calcium is important since a compound of calcium, lime, is used to make cement.

Magnesium is very light and it is used in aircraft and missiles since it has a light weight, but good strength when alloyed with other metals. Asbestos is a magnesium silicate.

O. Radioactive Metals

All metals, whose atomic number is larger than 83, are radioactive, along with many other elements in the periodic table. Being radioactive implies that the nucleus falls apart as it is not stable. On the periodic table, all elements on the bottom row are radioactive. Any new metals which would be produced in the laboratory would be radioactive, and they may only exist for a fraction of a second. Nuclear energy, as from uranium, may become an increasingly more important energy source in the future.

Very briefly, nuclear energy results because the heavy, radioactive nuclei fall apart, producing lighter elements. It takes much more energy to hold a heavy nucleus together than a nucleus from an atom in the middle of the periodic table. When uranium 235 breaks up, iron eventually forms. The mass of the original reactant, uranium, and the final products are almost the same, but not quite. The small amount of mass that is lost is converted to a huge amount of energy. In chemical reactions, mass is conserved, but this is a nuclear reaction, in which mass is converted into energy.

P. Resources in Alberta

Alberta is rich in many natural resources but not metals. Oil and gas fields are found throughout Alberta in abundance. Calcium-magnesium brine fields are found in central and southeastern Alberta. The foothills of the Rocky Mountains contain large limestone, dolomite and phosphate deposits.

There are small deposits of copper, lead, zinc, magnetite and titanium in the southwestern corner of Alberta. Also there are some low grade iron deposits west of Peace River. In the oil sands around Fort McMurray are deposits of cadmium, nickel, titanium and zirconium.

EXERCISES TO BE SENT FOR CORRECTION

Exercise 1 Alloys

1. Is it possible to get an aluminum-lead alloy? Why?

2. Give two properties which are often increased with alloying elements.

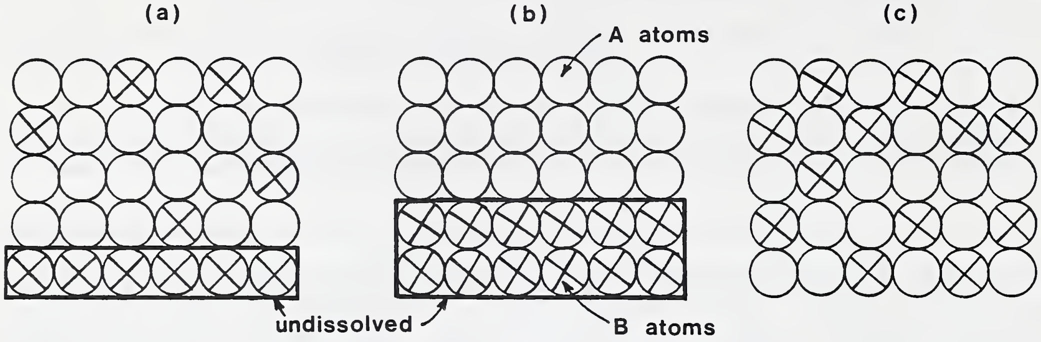
3. Give one property which is commonly decreased with alloying elements.

4. Which two metals are used in soldering?

5. Which two metals besides iron are major constituents of stainless steels?

6. With which two groups of metals are the salts (combination of metals with nonmetals) more important than other alloys?

7. In forming alloys, the amount of one metal which will dissolve another metal is often limited. If metal A could dissolve only 20% of metal B, which diagram illustrates what would happen if you mixed 40% of B with 60% of A?



Exercise 2 Noble Metals

1. What is an important feature of noble metals?

2. Name four noble metals.

3. Which two metals are most malleable?

Exercise 3 Refractory Metals

1. What is a common feature about refractory metals?

2. Name four refractory metals which are often alloyed with steel.

3. Give one reason why these refractory metals are often alloyed with steel.

Exercise 4 Miscellaneous Questions about various Metals

1. Which property of copper, that it shares with aluminum and silver, makes it important?

2. In conduction, which particles are most important?

   (a) protons
   (b) neutrons
   (c) inner electrons
   (d) outer electrons

3. Which property of aluminum makes it very good for aircraft parts?

4. Which three light metals are often used with aircraft?

5. Rearrange the following metals by their melting point, giving the one with the lowest melting point first. lead, tungsten, mercury, iron.

6. Which metal composes 99% of a "tin" can?

7. If you had a highly radioactive metal which was twice as strong as iron, would you use it in building constructions? Why or why not?

8. Most Uranium atoms have an atomic mass of 238. An atom of U^{238}

has how many protons ________, electrons ________ and neutrons ________.

9. Uranium 235, which is used for energy, has how many protons ________,

electrons ________, and neutrons ________. (Neutrons = Atomic Mass - Atomic Number)

10. If 3 g of tetraethyl lead are added for every gallon of gasoline, how many grams of pure lead would you be putting into your car if you filled it up with 10 gallons of gasoline?

Formula for tetraethyl lead is $Pb(C_2H_4)_4$, with a molar mass of 319 g. (Pb = 207.2 g)

11. We discussed two poisonous metals in greater detail, but there are many other poisonous metals. Beside each of the following poisonous metals give the symbol:

lead ______________ mercury ______________
nickel ______________ cadmium ______________
beryllium ______________ arsenic ______________

With these, and many other metals, particular compounds containing these metals are often much more dangerous than other compounds. Also, dangerous doses will vary from metal to metal. Many metals which are essential to human nutrition in small amounts may be deadly in larger amounts. Other metals, such as potassium and sodium, have to be in a certain balance with each other, so you shouldn't use an excess amount of table salt (NaCl) without making sure that you get enough potassium in your diet.

Exercise 5 Problems on Pollution

1. Quantities of lead in the atmosphere are often measured in micrograms per cubic metre of air. ($\mu g/m^3$). If the concentration of lead in the atmosphere of a city on a particular day is 1 $\mu g/m^3$, it means that there are 1×10^{-6} g of lead in a cubic metre of city air, since 1 g equals 1×10^6 μg. By inhaling the above air, assume a traffic policeman inhales a total of 20 μg (20×10^{-6} g) of lead in a 24 h period,

(a) calculate, in grams, the amount of lead inhaled in 2 h.

(b) Using proper scientific notation, calculate the number of moles of lead that are absorbed in the 2 h period. (molar mass of lead equals 207.2 g).

(c) Calculate the number of atoms that are absorbed in the 2 h period. (Hint: One mole of atoms equals 6.02×10^{23} atoms)

2. Lead poisoning occurs when the amount of lead in the blood reaches 0.8 ppm (ie. 0.8 g of lead in 1 million g of blood). Assuming that a policeman on patrol duty weighs 100 kg and has 8 kg of blood in his system, what would be the minimum amount of lead in his blood to indicate that he suffers from lead poisoning"

(Hint: Set up a ratio, $\frac{0.8\text{ g}}{1\ 000\ 000\text{ g}} = \frac{x\text{ g}}{8000\text{g}}$

Solve for x.

3. Pollution of air by industry and automobile emission is of a major concern to environmentalists. Suggest two or three ways that this type of pollution could be reduced.

1. ______________________________

____

____

____

2. ______________________________

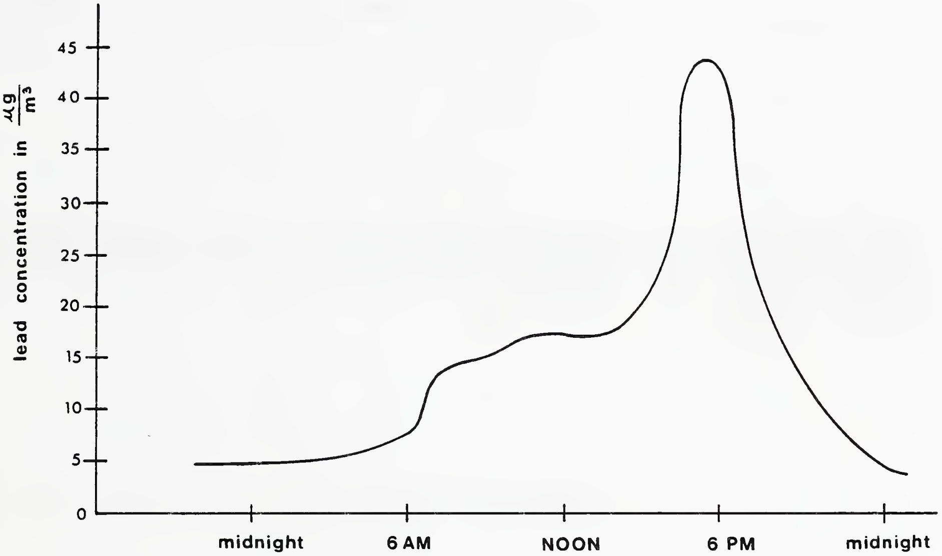
____

____

____

4. Acid rain, another serious pollutant is produced when sulfur dioxide from the air reacts with water to produce sulfurous acid (H_2SO_3). Write a balanced chemical equation for this reaction.

5. Following is a graph of lead concentration in the air of a large city, in the heart of the city, during a 24 h period.



Classify the following three observations as either qualitative or quantitative.

(a) The concentration increases as more cars are on the road. ________________

(b) During the afternoon rush hour period, the concentration reaches 45 $\mu g/m^3$. ________________

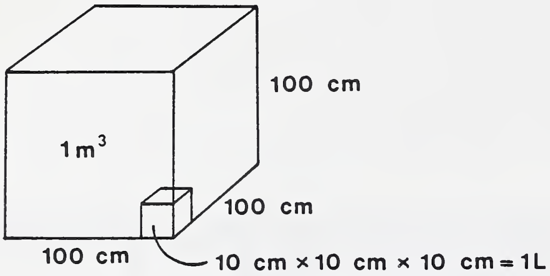
(c) There is always some lead in the air, regardless what time it is. ________________

(d) Is the statement in (a) an observation or an interpretation? ________________

(e) What is the concentration of lead at 3:00 p.m.? ________________

(f) Assume that when you take a deep breath of air, that you can inhale three litres of air. How many litres are in a cubic metre? ____________________

Hint



(g) How much lead would be in a litre of air at 3:00 p.m.? Give the answer in μg, and in grams, using exponential numbers for the grams. ($1\ \mu g = 1.0 \times 10^{-6}$ g)

(h) If you took a deep breath of 3 L, how much lead would you inhale in one breath?

(i) If 40% of the inhaled lead is absorbed by the blood, how much lead would this be for each big breath?

6. The Food and Drug Administration allows fish on the shelves to have a mercury concentration of 0.5 ppm. Assume that you bought a 200 g can of tuna fish with this concentration of mercury.

(a) What is the mass of mercury in this can?

(b) How many moles of mercury are there?

(c) How many atoms of mercury are there?

(d) Assuming that you weigh 70 kg, what would be the increase in mercury concentration in ppm in your body after you ate and digested the can of tuna?

(e) Assume that you ate a can of tuna for 21 days in a row with the same mercury concentration of 0.5 ppm. Also assume that your body was able to expel, by natural means, only one days supply of mercury in this time. How much higher would your concentration of mercury be in the body, in ppm, after 21 days?

Exercise 5 Alberta Resources

1. Name two metals found in the south western corner of Alberta.

2. Name two metals found in the oil sands around Fort McMurray.

3. Which metal is found west of Peace River?

4. Central Alberta and South eastern Alberta contain large amounts of which two metals?

Project

Metals have played an important part in the history of mankind, and in our modern society we have a large number of metals and alloys. The oil crises has forced us to try and think of alternative forms of energy, similarly, various metals are being depleted, and new metals or other substances may have to be found as an alternative. To do the project below it may be helpful to compare your house or car with a friend's house or car, or anything else, especially if there is a great difference in age of the house or car.

Find six or more objects or parts of objects for which one metal was used at one time, but for which a different metal is now used. List the metals used, or if it is an alloy, list the most prominent metal.

Object	Metal #1	Metal #2
1. cutlery	silver	steel
2.		
3.		
4.		
5.		
6.		
7.		

Along similar lines list six more objects in which a non metal is now used, where a metal was used in the past.

Object	Metal	Non Metal
1. top of salt shaker	stainless steel	plastic
2.		
3.		
4.		
5.		
6.		
7.		

END OF LESSON 11

LESSON RECORD FORM

1240 Chemistry 10

Revised 90/06

FOR STUDENT USE ONLY

Date Lesson Submitted ____________________

Time Spent on Lesson ____________________

(If label is missing
or incorrect)

File Number ____________________

Lesson Number __________

FOR SCHOOL USE ONLY

Assigned Teacher: __________

Lesson Grading: __________

Additional Grading E/R/P Code: __________

Mark: __________

Graded by: __________

Assignment Code: __________

Date Lesson Received: __________

Lesson Recorded __________

Student's Questions and Comments

Apply Lesson Label Here

Name __________

Address __________

Postal Code __________

Please verify that preprinted label is for correct course and lesson.

Teacher's Comments:

____

Correspondence Teacher

St. Serv. 21-89

ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

(1) All pages are numbered and in order, and no paper clips or staples are used.
(2) All exercises are completed. If not, explain why.
(3) Your work has been re-read to ensure accuracy in spelling and lesson details.
(4) The Lesson Record Form is filled out and the correct lesson label is attached.
(5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

First Class

Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope. Correspondence lessons will travel faster if first-class postage is used.**

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

REVIEW LESSON

This is the final lesson in your Chemistry 10 course. This lesson emphasizes the core lessons 1 to 8. If you have difficulty with any questions, refer back to the lessons and textbook to refresh your memory.

The final exam is a *closed book* exam. A periodic table is provided with the exam. Make sure that you bring your calculator and an extra pen and pencil. When writing your exam do all of the easy questions first then tackle the more difficult questions. Multiple choice questions should all be done since you are not penalized for incorrect answers. If you cannot get the answer to a multiple choice question reduce the choices as much as possible and then make an educated guess from the remaining choices.

We wish you success in your exam and your future endeavors.

Exercise 1 Fill in the blanks

2.5 km = ________________ cm

An area of 25 square m = ________________ square cm

500 cm^3 = ________________ L

75 km = ________________ m

30 m = ________________ cm

200 g = ________________ kg

15 cm = ________________ mm

800 mL = ________________ L

Exercise 2

An unknown metal whose mass was 85 g was placed into a 100 cm^3 graduated cylinder which was partially filled with water.

The height of water in the cylinder before emersion was 56.0 cm^3. Height after emersion was 75.8 cm^3.

Using the above data calculate the density of the unknown metal.

Exercise 3

Classify the following as either homogeneous or heterogeneous.

A. Honey

B. Oil and vinegar salad dressing

C. A sample of clean air

D. A teaspoon of baking powder

E. A bottle of homogenized milk left standing in a refrigerator for one day

F. A mixture of ice cubes and soda water

G. A mixture of undissolved sugar in a gallon of gasoline

Place the appropriate letters in the space below.

Homogeneous	Heterogeneous
____________________	____________________
____________________	____________________
____________________	____________________
____________________	____________________
____________________	____________________

Exercise 4

State whether the following items are metallic or non-metallic.

A. Insulation used in the walls of your home. ____________________

B. An arborite desk top. ____________________

C. A diamond in a wedding ring. ____________________

D. A piece of aluminum foil wrap. ____________________

E. A brass door knob. ____________________

F. A television antenna. ____________________

G. Floor tiles found in the kitchen of a home. ____________________

Exercise 5

A substance that dissolves in water to produce ions is called a(n) ____________________ solution.

List one example of this type of solution. ____________________

Exercise 6

Classify the following substances as elements or compounds. Use the letter E for an element and C for a compound in the space provided.

A. sulfuric acid ____________

B. tungsten ____________

C. natural gas ____________

D. vinegar ____________

E. gold ____________

F. helium gas ____________

G. chromium ____________

H. starch ____________

I. propane ____________

J. carbon ____________

Exercise 7

List four common examples where there is evidence of a chemical reaction.

E.g. 1. Natural gas burns in our hot water heater.

2. Drano is used to unplug a drain.

3. ____

____

4. ____

____

5. ____

____

6. ____

____

Exercise 8

Which of the following are observations and which are interpretations.

A. Ice melts at 0°C and sugar at 112°C. ____________

B. Heat and light are produced by a burning candle. ____________

C. The plateau on the cooling curve of a liquid indicates a phase change. ____________

D. All substances do not have the same melting points. ____________

E. Wilhelm Roentgen believed that the glow of the fluorescent material in the jar (see diagram p. 138 textbook) was caused by the same rays that were produced in the cathode-ray tube. ____________

Exercise 9

Classify the following properties as either physical (P) or chemical (C).

A. Molten wax drips along the edge of a candle and then solidifies. ____________

B. A slice of bread is toasted. ____________

C. A bottle of milk turns sour in a refrigerator. ____________

D. Aluminum melts at 658.7°C. ____________

E. Automobile fenders show corrosion after a year of driving. ____________

F. A sample of water sweetened with sugar is evaporated at room temperature whereby a white solid residue is obtained. ____________

Exercise 10

Give two examples of potential energy.

1. ____________

2. ____________

Exercise 11

Matter can be separated by physical or chemical means. After each description state whether separation involves physical or chemical means.

A. Gold nuggets are mixed with saw dust. The saw dust is burned off to leave only the gold. ____________

B. Gold and saw dust are placed into a large trough filled with water. The saw dust floats on water and is skimmed off. ____________

C. A mixture of alcohol and water is placed into a condenser in which alcohol is allowed to boil off at 78°C and recondensed in a cooling chamber. ____________

D. Liquid air is separated into its component gases by boiling off the various liquid components at different temperatures. ____________

E. An oil slick was set on fire to prevent pollution of the environment. ____________

F. Iron filings were accidentally mixed with gold dust. A chemist used some dilute hydrochloric acid to dissolve the iron. ____________

G. A sugar-coated spoon is placed into water, whereby sugar disappears from the spoon. ____________

Exercise 12

You have studied about several families of elements in this course. After each description below, place the name of the appropriate family in the space provided. Families to choose from are:

A. Noble gases

B. Alkali metals

C. Halogens

D. Alkaline earth metals

1. Molecules are diatomic. ____________

2. Have one electron in the outermost energy level. ____________

3. Have the greatest atomic radius. ____________

4. Each member contains one less electron than the following noble gas. ____________

5. Electron arrangements are very stable. ______________

6. Melting points increase as atomic numbers increase. ______________

7. Belong to group IIA. ______________

8. These elements form ions with a 1+ charge. ______________

9. These elements gain one electron to form negatively charged ions. ______________

10. These elements have two outer electrons. ______________

Exercise 13

Fill in the blanks.

An atom with an atomic number of 13 and an atomic weight of 27 has __________ protons, _________ electrons and __________ neutrons. It would be found in Group _____ of the periodic table and would be classified as a metal or nonmetal (underline). It would have _______ electrons in its outermost shell.

A sodium ion has _______ electrons and _______ protons giving it an overall charge of _________.

Exercise 14

List very briefly what each of the following scientists contributed to the field of science.

A. Avogadro ____

____

____

B. Rutherford ____

____

____

C. Dalton

D. Bohr

E. Chadwick

F. Mendeleev

G. Berzelius

Exercise 15

Name the following compounds.

A. $LiOH$

B. $Ba_3(PO_4)_2$

C. NH_3

D. $FeSO_4$

E. $Hg(ClO_3)_2$

F. NO_3

G. $SnCl_2$

H. $NH_4C_2H_3O_2$

I. KI

J. HNO_3

Exercise 16

Write the formula for each of the following.

A. sodium chlorite ______

B. iron III nitrate ______

C. tin IV chlorate ______

D. barium peroxide ______

E. tin II oxide ______

F. sodium bicarbonate (sodium hydrogen carbonate) ______

G. magnesium nitride ______

H. mercury I bromide ______

I. mercury II bromate ______

J. potassium permanganate ______

Exercise 17

Calculate the molecular weight of the following air pollutants.

A. SO_2 ______

B. H_2S ______

C. CCl_2F_2 (freon) ______

D. $PbBr_2$ ______

Exercise 18

1. Calculate the number of moles of carbon dioxide, CO_2, in 8.8 kg of compound.

2. Calculate the number of kilograms in 70 mol of aluminum sulfate, $Al_2(SO_4)_3$.

Exercise 19

Consider 120 g of acetic acid, CH_3COOH, in questions 1 to 6.

1. How many moles of acetic acid are present?

2. How many moles of carbon atoms are present?

3. How many moles of hydrogen atoms are present?

4. How many molecules of acetic acid are present?

5. What is the total number of moles of atoms present?

6. How many atoms of oxygen are present?

Exercise 20

Equations were classified according to four types of reactions (Lesson 8)

What type of reaction does each of the following equations represent?

A. $2Cu + O_2 \rightarrow 2CuO$ ______________________
(Copper pans tarnish in the air)

B. $2H_2O_2 \rightarrow 2H_2O + O_2$ ______________________
(hydrogen peroxide deteriorates when left in your medicine cabinet over a long period of time)

C. $C_2F_4 + C_2F_4 \rightarrow C_4F_8$ ______________________
(production of Teflon)

D. $NaHCO_3 + HCl \rightarrow CO_2 + H_2O + NaCl$ ______________________
(Treating acid burns)

E. $Cu + 2AgNO_3 \rightarrow 2Ag + Cu(NO_3)_2$ ______________________

Exercise 21

The incomplete burning of gasoline (C_8H_{18}) in the automobile engine produces a major air pollutant, carbon monoxide. The other product of combustion is steam ($H_2O_{(g)}$). Write a balanced equation to show the incomplete combustion of gasoline. (NOTE: Combustion requires oxygen from the atmosphere and also assume that no CO_2 is produced)

Exercise 22

Why should one not operate a vehicle in a closed or poorly ventilated garage?

____

____

____

Exercise 23

Water undergoes electrolysis to produce hydrogen and oxygen gas according to the equation below.

$$H_2O(\ell) + 285\ \text{kJ} \rightarrow H_2(g) + 0.50O_2(g)$$

From the above equation you will note that it takes approximately 285 kJ in terms of electrical energy to decompose one mole of water into its component elements.

Calculate the amount of energy required if 45 g of water undergo electrolysis.

Hint: First calculate the number of moles in 45 g of water, then proceed accordingly.

END OF LESSON 12